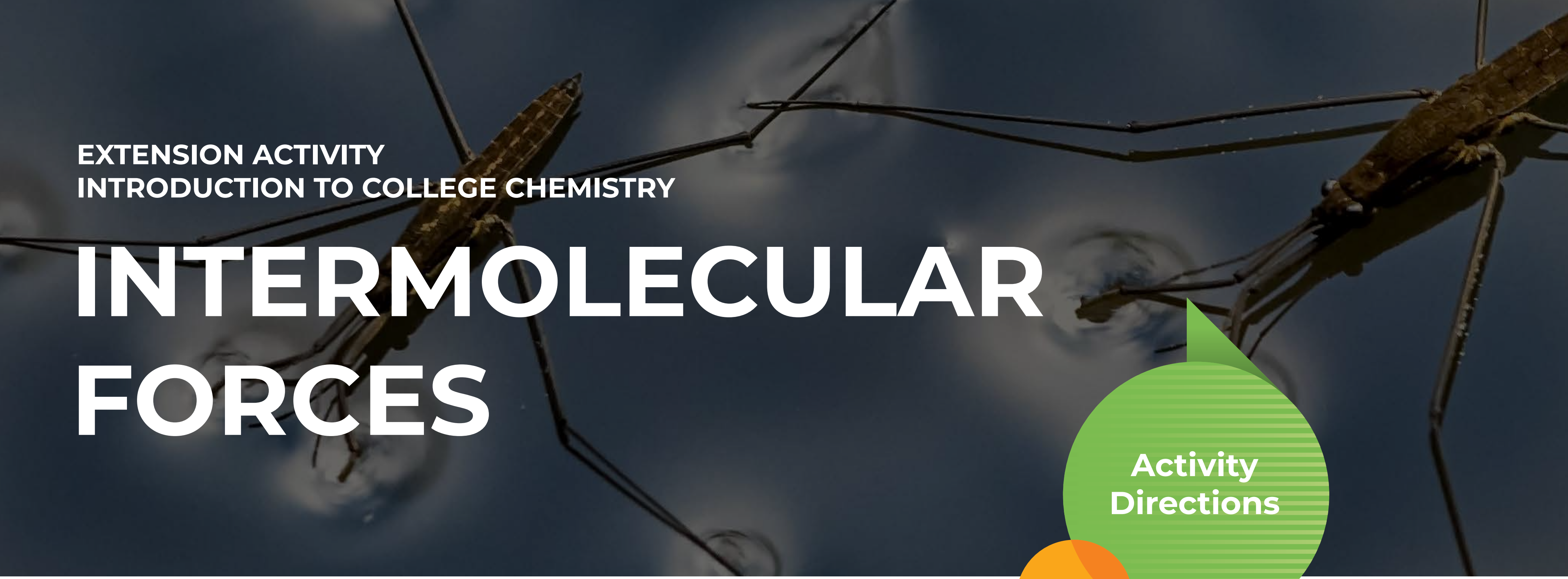


EXTENSION ACTIVITY
INTRODUCTION TO COLLEGE CHEMISTRY

INTERMOLECULAR FORCES

This activity will serve as practice for the topics covered in the Intermolecular Forces game. This activity is best used in conjunction with not only the tutorial levels, but also supplementary learning resources such as course lectures, textbook reading, etc. Questions labeled “Lock It In” are simply opportunities for you to solidify what you have accomplished in each task and help ensure you meet each objective.

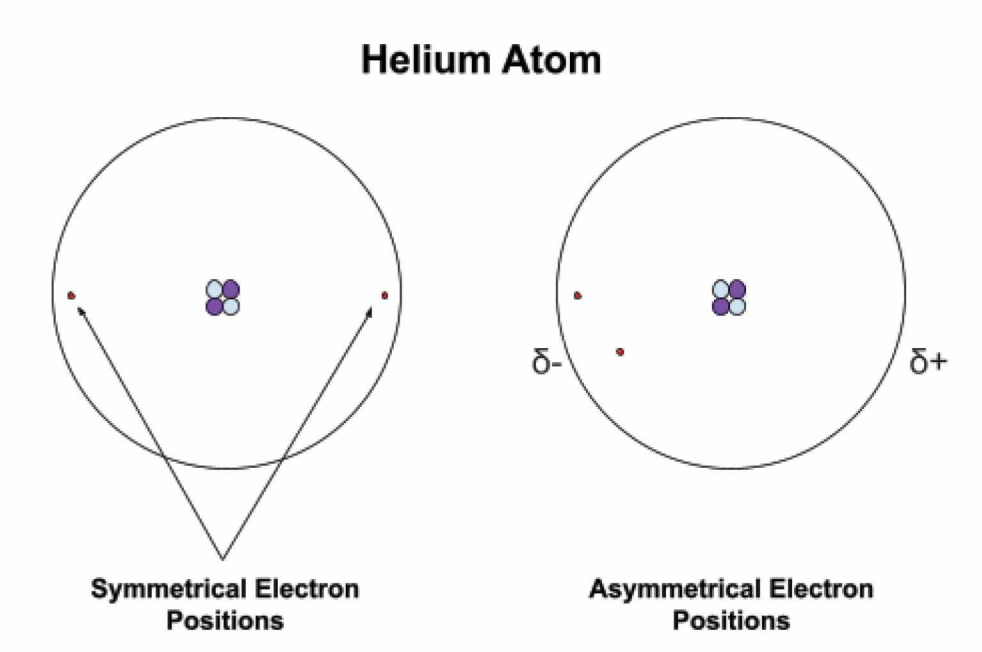
Activity Directions

1. Log into Collisions and navigate to the Intermolecular Forces Game.
2. Play the Tutorial levels, if you haven’t done so already.
3. Exit the levels and enter the Intermolecular Forces sandbox.
4. Follow all instructions as written below. Be sure to reference your course’s textbook, lecture notes, etc. as needed.

OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.

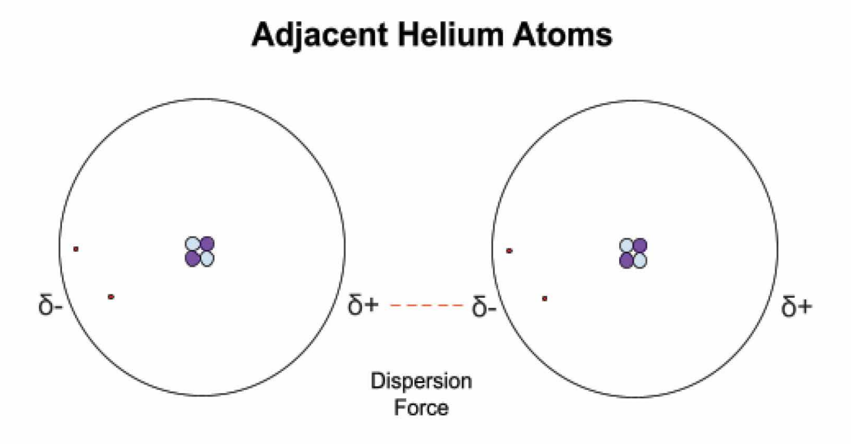
One important concept to remember in chemistry is that atoms, molecules, and their components are always in motion. This motion gives rise to very important phenomena, one of which is demonstrated by London dispersion forces. Although the Bohr molecule of the atom which shows electrons moving around the nucleus on fixed orbits is very commonly encountered, it is certainly not the most accurate (see the Radii Trends Game). Electrons are constantly making largely unpredictable movements within certain regions around the atom forming an electron cloud. As a result, there are moments when, by chance, electrons become more concentrated in one region of the electron cloud than others. In the image below you can see that in the instant captured on the left, the electrons of the helium atom are symmetrically spread out. However, in the instant captured on the right, you can see that the electrons have happened to be in the same region. Since electrons are negatively charged, this excess of electrons in one region combined with a dearth of them in the other creates an area of negative and positive charge respectively. The momentary separation of charges is called an **instantaneous dipole**. The temporary nature of this dipole is in contrast to the permanent dipoles of polar molecules.



OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.

As shown in the image to the right, the instantaneous dipole of one atom or molecule can create an instantaneous dipole on its neighboring atoms or molecules. These atoms or molecules will then attract one another electrostatically. This attraction is called the **dispersion force**, or **London dispersion force**.



The dispersion force will be the strongest intermolecular force in substances that do not have permanent dipoles. Although dispersion forces can be quite substantial as molecules increase in their number of electrons, permanent dipoles generate stronger forces in many small molecules. An easy way to think of this is that a substance composed of nonpolar molecules will be **LIMITED TO** dispersion forces, while a substance composed of polar molecules will have other forces in addition to dispersion forces.

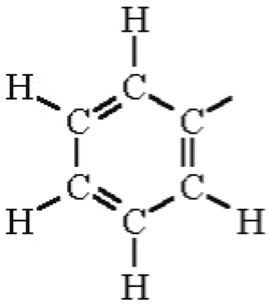
Dispersion forces are the predominant IMF in **monatomic** substances (substances with only one atom like the noble gases helium, argon, etc.) and **homonuclear** molecules (molecules containing only one element like nitrogen gas N_2). However, IMFs in **compounds** (substances with more than one element) will vary depending on composition and geometry. Please note the word “predominant” because ALL substances will have London dispersion forces, while some molecules will have additional intermolecular forces present.

It is important to note that the fleeting nature of the instantaneous dipole makes individual dispersion forces quite weak. However, there is considerable variation in the strength of dispersion forces based on how well an atom or molecule is able to polarize (form a dipole). This process is more easily accomplished in atoms or molecules with larger electron clouds. As a result, the molar mass of a substance is a good (but not perfect) proxy for comparing the strength of its London dispersion forces to those of another substance. In other words, increasing size typically leads to increasing dispersion forces. As such, a large molecule like pentacontane ($C_{50}H_{102}$) will have far larger dispersion forces between molecules than that of the much smaller butane (C_4H_{10}).

OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.

TASK 1: To explore the concept of LDFs a little more, let's take a look at parts of two different molecules. Below you will see what are called the **side chains** of the amino acids tyrosine and phenylalanine. The side chain is the portion of an amino acid that distinguishes it from other amino acids and gives it a lot of its character. Different side chains have different capabilities partially due to their different polarities. As such, they present a good opportunity for us to practice identifying side chains in which London dispersion forces will predominate in its intermolecular interactions versus those that will have other interactions. Below each side chain you will see a table that identifies the different bonded elements in the structure. Classify each of the bonds in the table below as polar or nonpolar using their electronegativities provided on page 6.

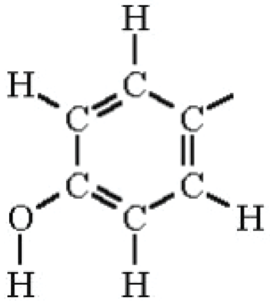


Phenylalanine Side Chain

Chemical Bond	Electronegativity of First Element	Electronegativity of Second Element	Electronegativity Difference	Polar or Nonpolar?
C-C or C=C				
C-H				

OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.



Tyrosine Side Chain

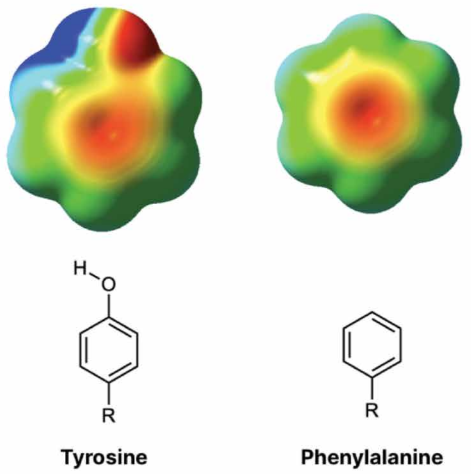
Chemical Bond	Electronegativity of First Element	Electronegativity of Second Element	Electronegativity Difference	Polar or Nonpolar?
C-C or C=C				
C-H				
C-O				
O-H				

LOCK IT IN:
For which of the two side chains shown above will London Dispersion Forces be the only intermolecular force in its interactions? Explain your answer.

OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.

Below are the **electrostatic potential maps** for each side chain. On the electrostatic potential maps below, bright red indicates areas of the molecule that have a more partially negative charge due to electrons spending more time in those locations. Dark blue indicates areas that have a more partially positive charge due to electrons spending relatively less time around them. Other colors represent more intermediate tendencies to have shared electrons in those areas.

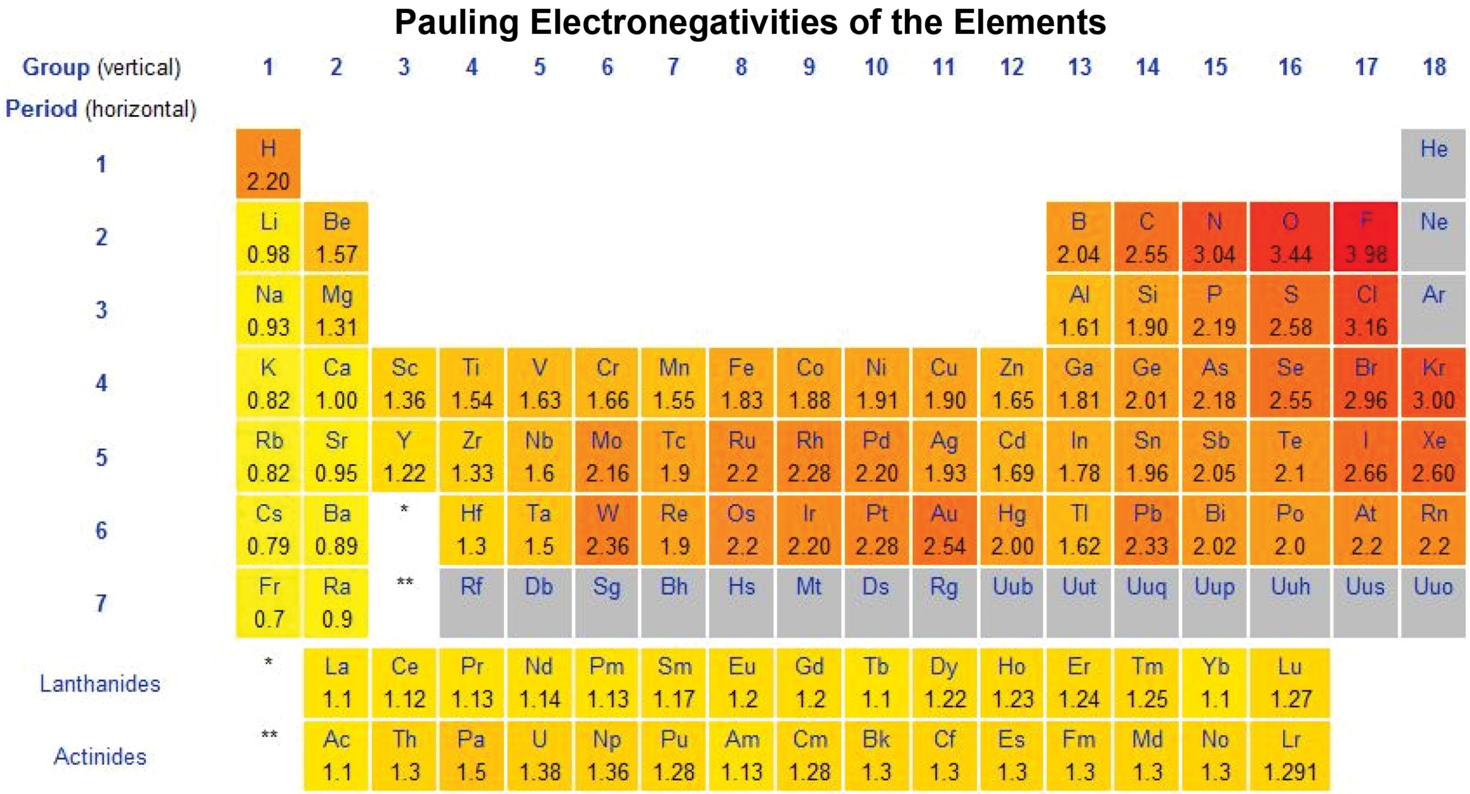


Electrostatic Potential Maps of Aromatic Amino Acids, Emanresucamit CC BY-SA 4.0

LOCK IT IN:
Do the electrostatic potential maps confirm your decision as to which side chain will have dispersion forces as its only intermolecular force? Why or why not?

OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.



Pauling Electronegativities of the Elements

Group (vertical)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Period (horizontal)																		
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo
Lanthanides	*	La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27		
Actinides	**	Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.291		

Electronegativity Values, CC BY-SA 3.0

Electronegativity Difference	Bond Type
< 0.4	Nonpolar Covalent
0.4 to 1.7	Polar Covalent
> 1.7	Ionic if between metal and nonmetal, but polar covalent if not

OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.

LOCK IT IN:
Phosphorus sulfides are a family of compounds that contain different quantities of phosphorus and sulfur. They include compounds with formulas like P_4S_{10}, P_4S_3, P_4S_5, etc. Some of the bonds in each structure are between phosphorus atoms, some are between sulfur atoms, and some are between sulfur and phosphorus atoms. Explain why the strongest intermolecular force in a quantity of any of these phosphorus sulfides will be London dispersion forces.

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OBJECTIVE 1

Demonstrate why dispersion forces exist and how they vary amongst molecules.

TASK 2: Identify six substances created using only the elements found in the sandbox that have London Dispersion forces as their strongest intermolecular force. Confirm your answers using the sandbox. At least one substance should fall into each of the following categories: monatomic, homonuclear molecule, and a compound. Rank the substances by the strength of the London dispersion force. Include the molar mass of each substance you create. See your textbook or notes for help determining molar mass as necessary. Use the electronegativity table to help you when necessary.

	Chemical Formula	Molar Mass (g/mol)	Monatomic, Homonuclear, or Compound?
Strongest LDF			
Weakest LDF			

LOCK IT IN:
Explain what reasoning you used to create molecules only showing London dispersion forces and how you were able to rank the strength of those forces.

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OBJECTIVE 2

Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

Dipole-dipole interactions are what occur when the instantaneous dipoles of the London dispersion forces are instead permanent. Due to differing electronegativities amongst the elements, the electrons in compounds are very often unequally shared. Bonds between atoms where electrons are shared unequally are called **polar bonds**. Very often, these polar bonds are arranged in a way that creates a permanent dipole on a molecule overall. These dipoles will then cause the molecules bearing them to be attracted to one another in the typical electrostatic fashion. Do note, however, that these dipoles are considered partial charges (represented by δ^- or δ^+) and are not nearly as strong as ions in their attraction to one another. These interactions are not even as strong as the covalent bonds within the molecules themselves. Regardless, dipole-dipole interactions can explain many fascinating chemical trends and phenomena.

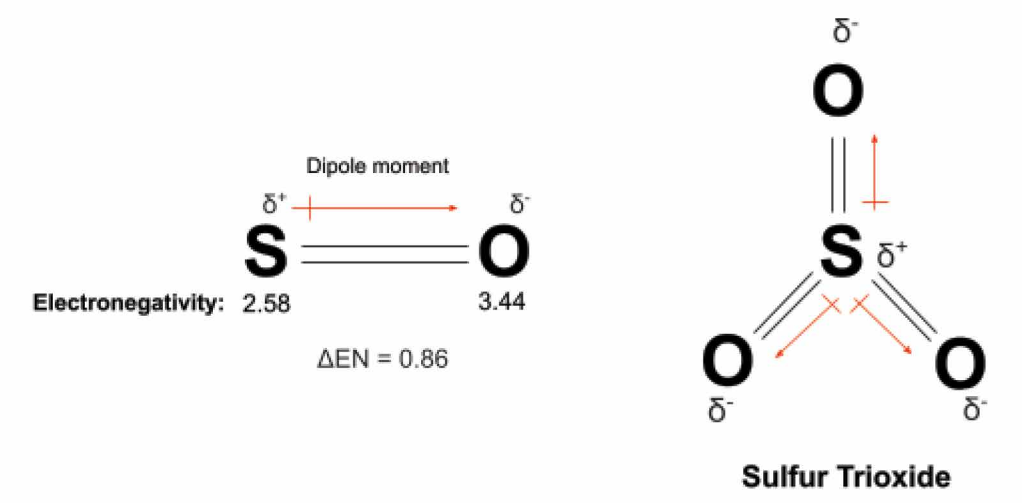
The permanent dipole in HCl due to electronegativity differences between hydrogen and chlorine creates the dipole-dipole interaction indicated by the dashed line in the image.

OBJECTIVE 2

Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

Interestingly enough, a molecule composed of only polar bonds can be considered nonpolar. This is because both the polarity of bonds within the molecule and its molecular geometry are necessary for determining the overall polarity of a molecule.

In order to understand how this works, one must understand the concept of a **bond dipole moment**. This concept is simply an extension of the bond polarity that you explored in Task 1. Bonded atoms often do not share electrons equally, and when this happens the bond is said to exhibit a dipole moment. This dipole moment can be represented by a vector (arrow) pointing from the less electronegative element to the more electronegative one. The now slightly more positive atom receives the label δ^+ while the slightly more negative one receives the label δ^-. Look at the examples with sulfur and oxygen below.

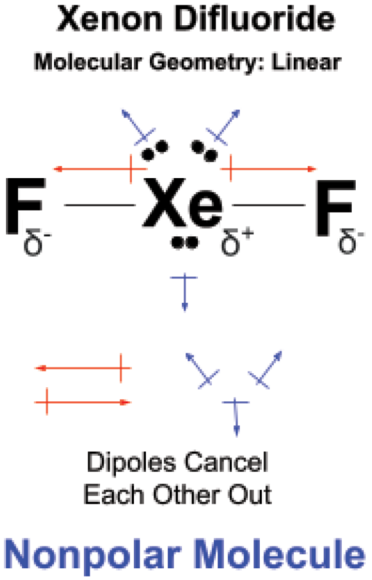


However, individual dipole moments can cancel each other out depending on the shape of the molecule. It is for this reason that it is important to determine the **net dipole moment** for the molecule by considering the sum of all individual dipole moments created by polar bonds. In fact, sulfur trioxide above is nonpolar despite having three polar bonds. Please reference your textbook or lecture notes for a deeper understanding of polarity before continuing.

OBJECTIVE 2

Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

Consider xenon difluoride and sulfur dioxide. Notice that both contain polar bonds, but only the molecular geometry of sulfur dioxide allows for a net dipole moment. As a result, the fully symmetrical xenon difluoride is a nonpolar molecule and sulfur dioxide is a polar one.

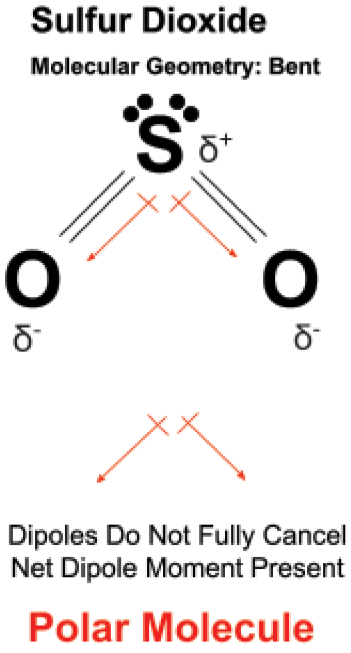


Xenon Difluoride

Molecular Geometry: Linear

Dipoles Cancel
Each Other Out

Nonpolar Molecule



Sulfur Dioxide

Molecular Geometry: Bent

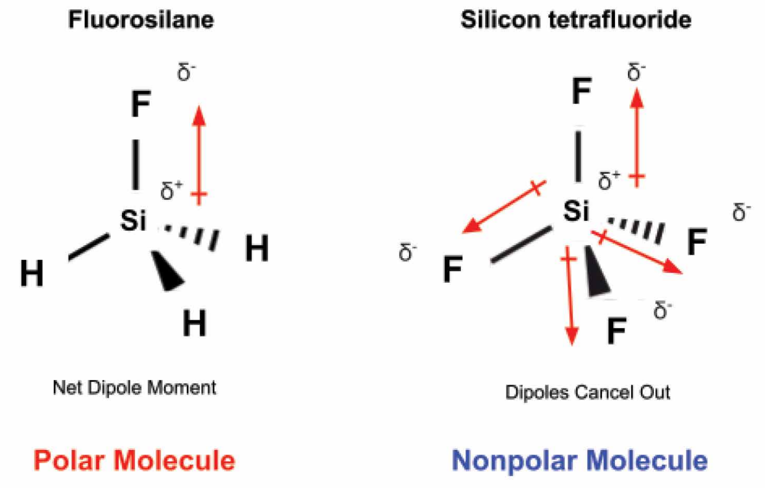
Dipoles Do Not Fully Cancel
Net Dipole Moment Present

Polar Molecule

OBJECTIVE 2

Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

Consider fluorosilane (SiH_3F). Although the molecular geometry of the structure (tetrahedral) is symmetrical, due to its composition it has one polar bond that is not canceled out by other equally polar bonds in the opposite directions. Compare it to silicon tetrafluoride (SiF_4) whose four polar bonds cancel each other out.



Determining whether a molecule is polar requires a few steps:

1. Draw a Lewis structure for the molecule.
2. Determine if the molecule has any polar bonds using electronegativities. A molecule without any polar bonds is **nonpolar**.
3. Determine the molecular geometry of the structure. If the atoms attached to the central atom are all the same, then molecules of the following geometries will be **nonpolar** due to their symmetry:
    - Linear
    - Trigonal Planar
    - Square Planar
    - Tetrahedral
    - Trigonal Bipyramidal
    - Octahedral
4. If the molecule has at least one polar bond and the atoms bonded to the central atom are not identical, the molecule is *probably* **polar**.
5. Assess the sum of the individual dipoles. If all polar bonds are countered by bonds equal in magnitude and opposite direction, then they will cancel out. Such molecules are **nonpolar**. In the case that a net dipole remains, the molecule is **polar**.

OBJECTIVE 2

Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

TASK 3: Classify each of the bonds in the table below as polar or nonpolar using their electronegativities. Then determine if dipole-dipole interactions would be **possible** in a molecule containing ONLY that bond. Use the table of electronegativities from Objective 1 to help you.

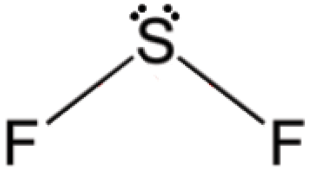
Chemical Bond	**Electronegativity of First Element**	**Electronegativity of Second Element**	**Electronegativity Difference**	**Polar or Nonpolar?**	**Possible contributor to dipole-dipole interaction? (Yes or No)**
C — N					
S — O					
Si — C					
I — Br					
P — O					

LOCK IT IN:
Chlorine monofluoride (CIF) only contains the bond between one chlorine atom and one fluorine atom. Explain how we know that chlorine monofluoride molecules will show dipole-dipole interactions.

OBJECTIVE 2

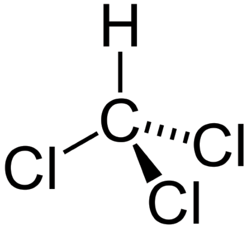
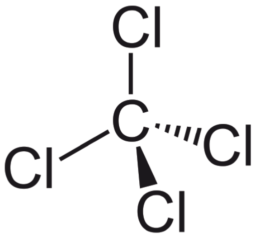
Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

TASK 4: Using what you know about polar bonds and molecular geometry, determine whether the molecules below will display dipole-dipole interactions. Explain your reasoning for each by referencing the bond polarities and molecular geometries for each molecule. Confirm your answers by constructing the structure in the sandbox. Please consult your textbook, lecture notes, and/or the Lewis Structure Extension Activity for more details on molecular geometry.

Chemical Name	Chemical Formula	Structure and Molecular Geometry	Dipole-Dipole Interaction? (Yes or No)	Justification
Hydrogen Cyanide	HCN	$H-C\equiv N$ Linear		
Carbon Dioxide	CO_2	$O=C=O$ Linear		
Sulfur Difluoride	SF_2	 Bent		

OBJECTIVE 2

Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

Chloroform	$CHCl_3$	 Tetrahedral		
Carbon Tetrachloride	CCl_4	 Tetrahedral		

LOCK IT IN:
Explain in your own words why molecules with polar bonds might not be polar molecules. Use one of the molecules from Task 4 to explain.

OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.

TASK 5: Use the electronegativity values from the periodic table under Objective 1 to rank the polarity of the bonds containing hydrogen below.

C — H O — H Br — H N — H F — H S — H P — H H — H

	Chemical Bond
Very Polar	
Nonpolar	

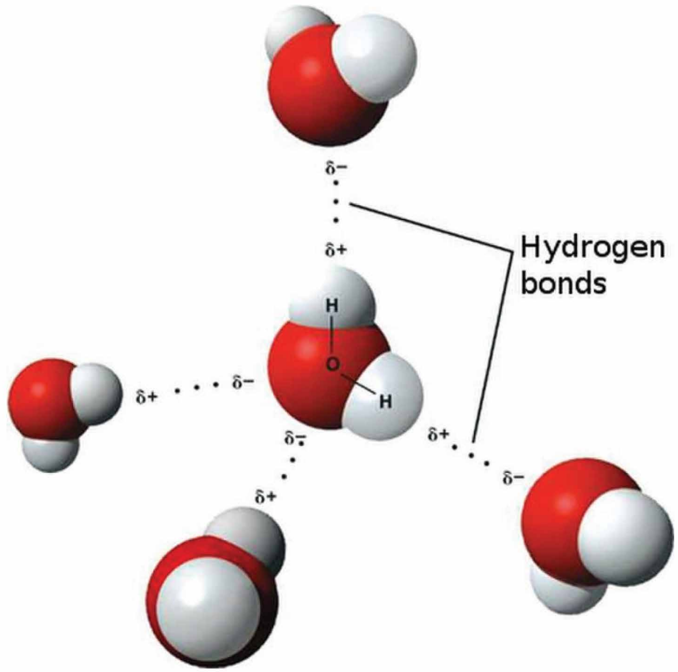
LOCK IT IN:
Which three elements from the list above produce the most polar bonds with hydrogen atoms?

OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.

As you should have seen in Task 5, strongly polar bonds are created when hydrogen atoms bond with very electronegative atoms (oxygen, nitrogen, and fluorine). When these bonds occur and the very electronegative atom holds at least one lone pair, a special dipole-dipole interaction known as a **hydrogen bond** forms between the partially positive hydrogen of one molecule and the partially negative atom of another. Water is one of the most famous molecules capable of hydrogen bonding and for good reason—it is extremely well-suited to it. There exist, however, many other examples.

Do note that hydrogen bonds are particularly strong compared to other intermolecular forces, but they are still weaker (less than a tenth of the strength) than the average covalent bond and much weaker than ionic ones. Regardless, hydrogen bonding is still responsible for a plethora of amazing occurrences in chemistry.



Hydrogen bonds depicted between water molecules.

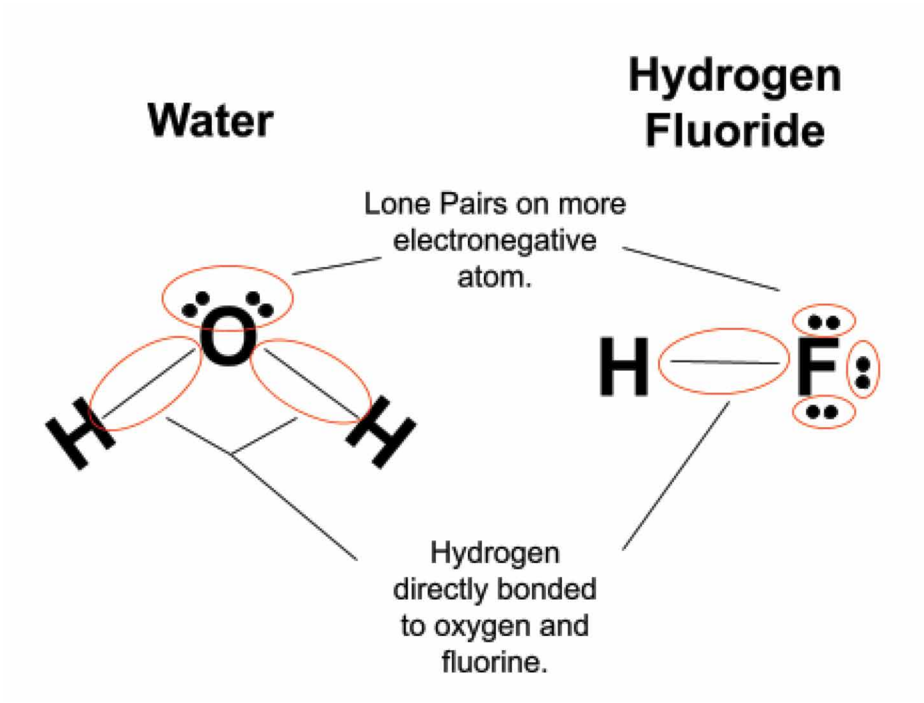
OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.

1. Determining if a Molecule is Capable of Hydrogen Bonding

To decide if a molecule is capable of hydrogen bonding, it helps to look at its molecular structure. Hydrogen must be (although there are some minor exceptions) attached directly to **oxygen, nitrogen**, or **fluorine**. Although chlorine is also very electronegative, it is generally considered too large of an atom relative to tiny hydrogen to create true hydrogen bonding. **The atom to which the hydrogen is attached must also have lone pairs.** The importance of the lone pairs will come up later.

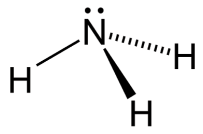
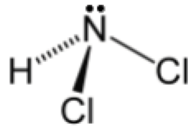
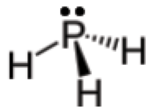
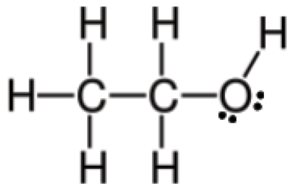
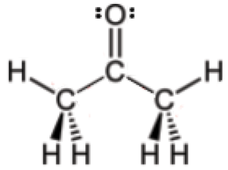
Look at the structures of water and hydrogen fluoride, both of which are capable of hydrogen bonding:



OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.

TASK 6: Determine if each molecule below would be able to form hydrogen bonds. To make sure you are on the right track, check your answers for the first two molecules using the sandbox.

Chemical Name	Chemical Formula	Structure	Hydrogen Bonding? (Yes or No)
Ammonia	NH_3		
Dichloramine	$NHCl_2$		
Phosphine	PH_3		
Ethanol	C_2H_6O		
Acetone	C_3H_6O		

OBJECTIVE 3 Demonstrate an understanding of why hydrogen bonding occurs.

LOCK IT IN:
Describe the method you used to determine whether each molecule was capable of hydrogen bonding.

OBJECTIVE 4

Demonstrate an understanding of the role of polarity in the solubility and miscibility of covalent substances.

Another very important concept in chemistry is that of **solubility**— the ability of a substance to dissolve in another and form a **solution**. A solution is composed of a **solvent** (the substance doing the dissolving) and a **solute** (the substance being dissolved). Intermolecular forces are so important to solubility because solvent-solvent, solute-solute, and solvent-solute interactions will determine just how much of a substance dissolves in another. As a result, chemists often like to say that **"like dissolves like"** or that substances often show high solubility in substances that have similar intermolecular forces to their own. Nonpolar substances will often form solutions with other nonpolar substances and polar substances will often form solutions with other polar substances. Hydrogen bonding capabilities, for example, can increase the solubility of a substance in water since the solvent and solute will be able to form hydrogen bonds with one another.

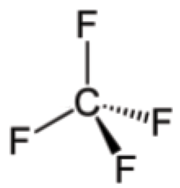
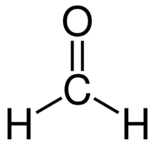
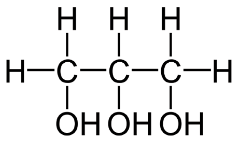
You have certainly noticed that oil and water do not mix. This is the result of them being insoluble in one another. A term more commonly used to describe the two is **immiscible. Miscibility** describes the ability of two substances to mix in all ratios due to their high solubilities in one another. For example, acetone (found in nail polish remover) is miscible in water. This means that one could make a solution that is 1% water and 99% acetone, 50% water and 50% acetone, or 99% water and 1% acetone. Solubilities are usually given as the maximum mass of a substance that can dissolve in a certain volume or mass of another. However, acetone and water are so soluble in one another that they mix freely and are simply described as miscible. Both acetone and water are polar molecules with significant, similar intermolecular interactions. Their miscibility makes sense. However, nonene, which contains only bonds between carbon and hydrogen, is very nonpolar and thus very insoluble in water.

Oil and Water, Victor Blacus CC BY-SA 4.0

OBJECTIVE 4

Demonstrate an understanding of the role of polarity in the solubility and miscibility of covalent substances.

TASK 7: Determine the strongest IMF that would be present in a quantity of each molecule listed below. Make the first four in the sandbox to check your answers. Note that use of “insoluble” indicates a solubility of less than 1 gram/Liter.

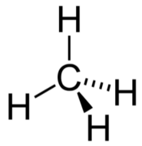
Chemical Name	Chemical Formula	Structure	Predominant IMF	Soluble in Water at 20°C and 1 atm?
Carbon Tetrafluoride	CF_4			insoluble
Formaldehyde	CH_2O			soluble
Neon	Ne	Ne		insoluble
Dioxygen	O_2	$O=O$		insoluble
Glycerol	$C_3H_8O_3$	 $\begin{array}{ccccccc} & H & & H & & H & \\ & \vert & & \vert & & \vert & \\ H- & C & - & C & - & C & -H \\ & \vert & & \vert & & \vert & \\ & OH & & OH & & OH & \end{array}$		miscible

OBJECTIVE 4

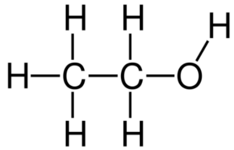
Demonstrate an understanding of the role of polarity in the solubility and miscibility of covalent substances.

LOCK IT IN:
Take note of the substances in the table that are most and least soluble in water. What types of intermolecular forces seem to increase the solubility of a substance in water? Why does that make sense?

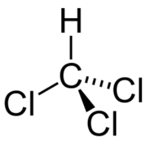
TASK 8: In the box are three substances that you have encountered as completing this extension activity. Use their structures to rank them in terms of their solubility in water and then provide a brief justification as to your ranking.



Methane
CH_4



Ethanol
C_2H_6O



Chloroform
$CHCl_3$

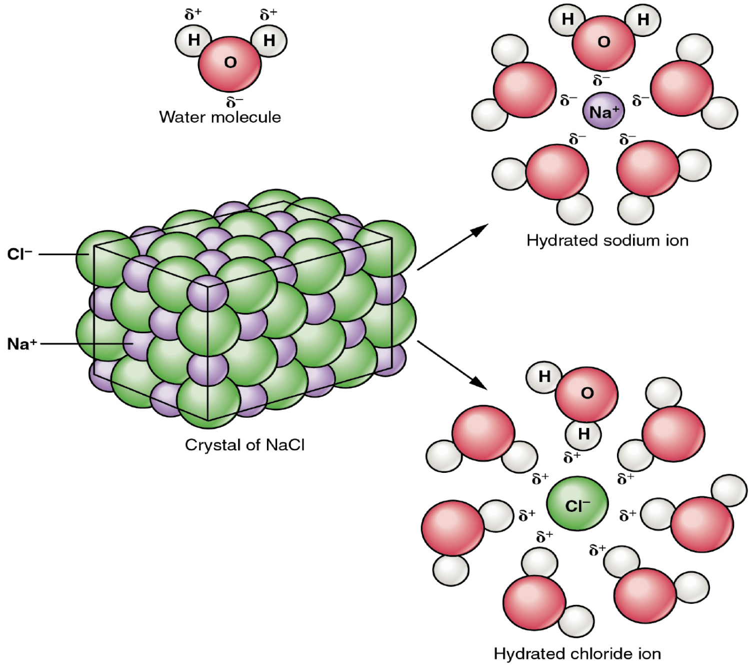
	Chemical Name and Formula	Justification
Most Soluble in Water		
Least Soluble in Water		

OBJECTIVE 5

Demonstrate an understanding of how ion-dipole interactions allow ionic substances to form solutions in water.

Another intermolecular force is known as an ion-dipole interaction. Ion-dipole forces are stronger than all of the other intermolecular forces covered in this extension activity, and they are a particularly important interaction when considering solutions made when ionic compounds dissolve in covalent ones.

As we know, ions are by definition either positively charged cations or negatively charged anions. The charges of these ions present interesting interactions with the partial charges of dipoles on polar molecules. If we continue to use the idea that “like dissolves like” as we saw when thinking about the solubilities of covalent substances, then it should be unsurprising to learn that ionic compounds (which are charged) are typically more soluble in polar solvents (which have partial charges).



NaCl dissolution, Open Stax College [CC BY 3.0]

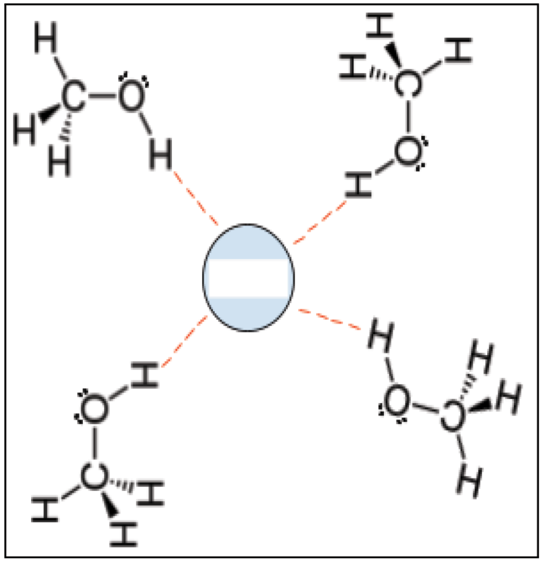
The lattice energy of ionic compounds is not easily overcome, but the dipoles of polar solvents are often capable of doing so because of how they arrange themselves around individual ions to create energetically favorable conditions within a solution. For example, the polarity of water molecules are what make water such a good solvent for ionic compounds.

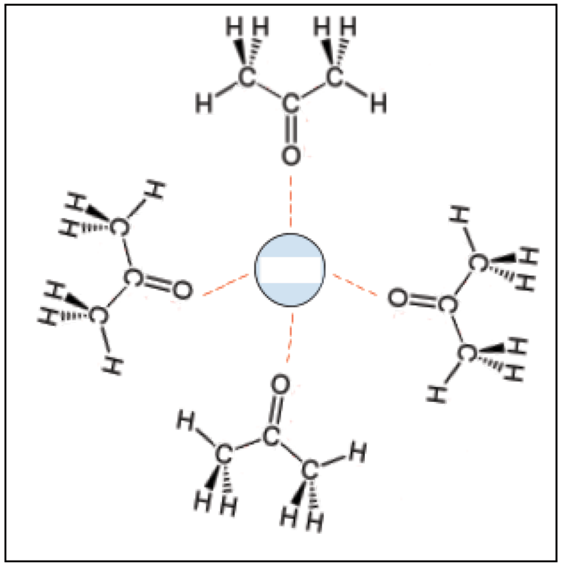
The image to the left depicts water molecules forming **hydration shells** around sodium and chloride ions that have been separated from one another. Notice how the oxygen atom bearing the partial negative charge is oriented towards the positively charged sodium ion, while the hydrogen atoms with the partial positive charge are oriented towards the negatively charged chlorine ion. Such an orientation of dipoles around ions is characteristic of ion-dipole interactions.

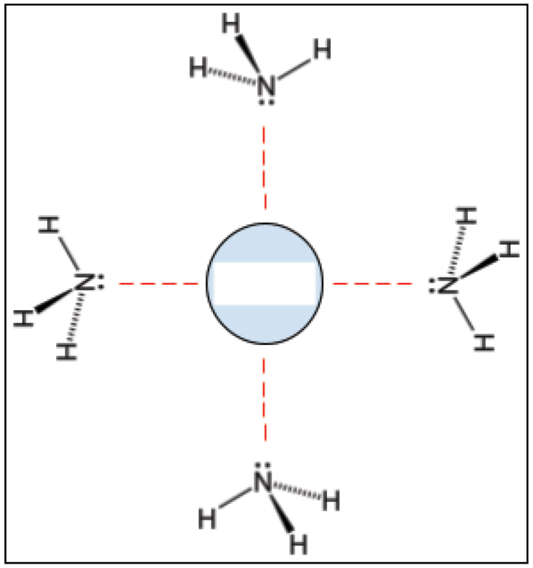
OBJECTIVE 5

Demonstrate an understanding of how ion-dipole interactions allow ionic substances to form solutions in water.

TASK 9: Label each ion-dipole interaction below by labeling the ion in the center of the image with a “—” for an anion or a “+” for a cation depending on the orientation of the surrounding molecules.



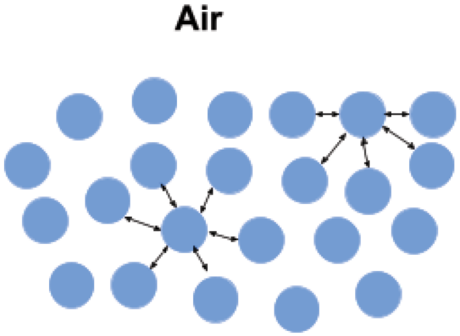




OBJECTIVE 6

Demonstrate an understanding of how intermolecular forces influence surface tension, viscosity, and capillary action.

Surface tension is defined as the amount of energy required to increase the surface area of a liquid by a certain amount. This property is very much a result of intermolecular forces. The simplest way to understand this concept is to consider a certain quantity of water that is interacting with air. Due to hydrogen bonding, a water molecule is going to experience significant intermolecular forces with other water molecules, but much less interaction with the molecules of the air. A water molecule away from the surface is going to experience hydrogen bonding with molecules all around it that lead to no net pull in any direction. However, a water molecule at the surface interacting with air will experience much greater pull towards the interior of the water than it will experience towards the air. This leads to a net pull inwards and results in water minimizing its surface area. The relatively large amount of energy required to overcome this tendency of water to reduce its surface area leads to water having a high surface tension of 0.075 Newtons/meter at 25°C.



The surface molecule experiences greater forces pulling it towards the interior than towards the air.

Paper Clip on Water, Armin Kübelbeck CC BY-SA 3.0

Water's high surface tension both allows the water in the glass to reach slightly higher than the rim without spilling and allows the paperclip to float on the surface. Many insects take advantage of this property.

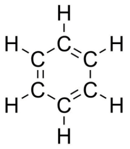
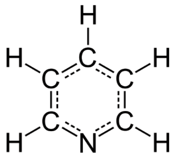
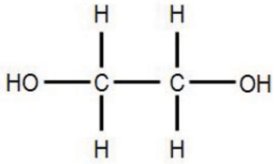
OBJECTIVE 6

Demonstrate an understanding of how intermolecular forces influence surface tension, viscosity, and capillary action.

Stronger intermolecular forces lead to a greater net pull towards the interior of a liquid, which means stronger intermolecular forces lead to higher surface tensions. As such, one can look at the structure and composition of a molecule to get an idea of how its surface tension would compare to that of another substance.

TASK 10: Listed in the box below are the surface tensions of three substances at 25°C given in N/m (Newtons per meter). Match them with the correct substance and provide your justification.

0.048 0.028 0.037

Chemical Name	Chemical Formula	Structure	Surface Tension (N/m)	Justification
Benzene	C_6H_6			
Pyridine	C_5H_5N			
Ethylene Glycol	$C_2H_6O_2$			

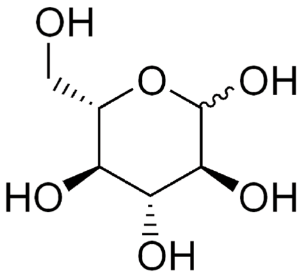
OBJECTIVE 6

Demonstrate an understanding of how intermolecular forces influence surface tension, viscosity, and capillary action.

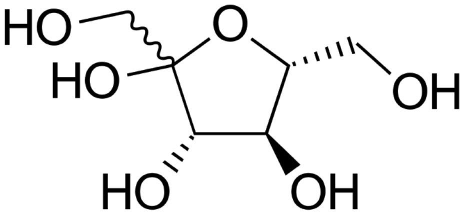
A key characteristic of liquids is their ability to flow. However, liquids like water will flow much more easily than liquids like honey. The term used to describe a liquid's resistance to flow is **viscosity**. As such, we say that honey is more viscous (has a higher viscosity) than water. Viscosity is the result of the intermolecular forces found in a substance. Greater intermolecular forces typically lead to higher viscosities. Larger dispersion forces, the presence of hydrogen bonding, and/or dipole-dipole interactions can all increase the viscosity of a substance.

Honey in a Glass, Lama Raheem CC BY-SA 4.0

Glucose



Fructose



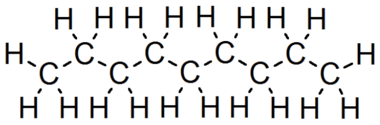
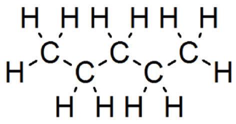
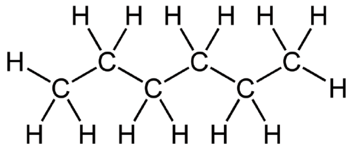
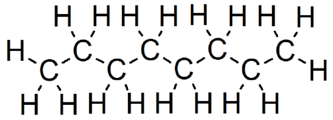
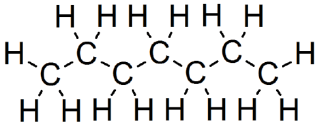
Hydrogen bonding between the sugar molecules of glucose and fructose help give honey its high viscosity.

OBJECTIVE 6

Demonstrate an understanding of how intermolecular forces influence surface tension, viscosity, and capillary action.

TASK 11: Listed in the box below are the viscosities of five substances at 25°C given in mPa·s (millipascal-second). Match the substances with their viscosities.

0.509 0.224 0.389 0.295 0.665

Chemical Name	Chemical Formula	Structure	Molar Mass (g/mol)	Viscosity (mPa·s)
n-Nonane	C_9H_{20}		128.3	
n-Pentane	C_5H_{12}		72.15	
n-Hexane	C_6H_{14}		86.17	
n-Octane	C_8H_{18}		114.2	
n-Heptane	C_7H_{16}		100.2	

OBJECTIVE 6

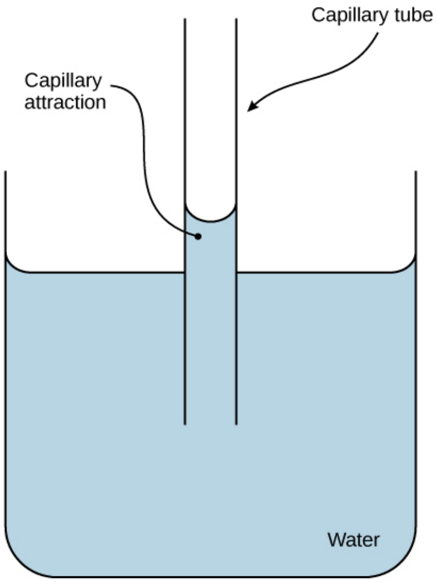
Demonstrate an understanding of how intermolecular forces influence surface tension, viscosity, and capillary action.

LOCK IT IN:
Explain the logic you used to determine which viscosity matched each molecule.

Intermolecular forces also play a crucial role in a concept known as **capillary action**—the ability of a liquid to move upwards against gravity in a narrow tube. You might have noticed as a child that when you place a straw in a drink, the liquid rises into the straw a bit higher than the surrounding liquid in the cup. This is capillary action. Although the straw is just a simple example, capillary action has many important implications, including helping water travel from the soil into and up plant stems.

Understanding capillary action requires thinking about two categories of intermolecular forces—**cohesive forces** and **adhesive forces**. Cohesive forces describe the interactions between molecules of the same type, while adhesive forces describe the interactions between different molecules. In order for capillary action to occur, the liquid must experience greater adhesive forces with the wall of the tube than it does cohesive forces with the other molecules of its type. As the liquid is attracted to the wall of the tube, molecules will move upwards against gravity while also pulling up neighboring molecules through cohesive forces.

When water enters a glass tube, the water molecules are attracted to the glass surface because it is polar. In fact, the water molecules are more attracted to the glass surface than to other water molecules. The result is that the water can move upward in the tube against gravity. Had the glass not been polar, then the adhesive forces between water and the glass would have been insufficient for capillary action to occur.



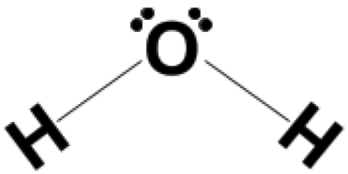
Capillary Action, Open Stax College CC BY 4.0

OBJECTIVE 6

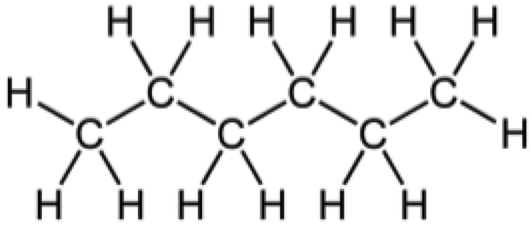
Demonstrate an understanding of how intermolecular forces influence surface tension, viscosity, and capillary action.

TASK 12: A thin, glass tube is placed into a beaker of n-hexane and another into a beaker of water. In the space below, compare the substances based on how far up the glass tube they will travel. Explain your reasoning using what you read above.

Water



Hexane



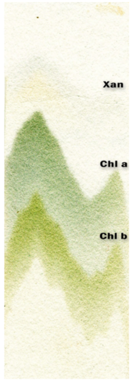
OBJECTIVE 7

Demonstrate an understanding of how chromatography can be used to separate the components of mixtures by taking advantage of variation in intermolecular forces between molecules.

One of the key lab techniques frequently used by chemists, biochemists, biologists and other scientists is **chromatography**. Chromatography is a technique used to separate mixtures of substances so that individual components can be analyzed. To help us understand how chromatography works, it might first help to imagine a trip to a carnival. You and a large group arrive at a carnival at the same time and begin to work your way from one side of the carnival to the other. But a group of people as large as yours will hardly make it to the other side at the same time. Everyone is different and there is considerable variation in the interests of the group. Some members of the group will quickly move through the carnival uninterested in the myriad vendors, rides, and distractions. These people will make it to the other side of the carnival fastest. Other members will stop at every vendor, hawker, and ride at the carnival. These people will be the last to arrive at the other side of the carnival. Other members of the group will fall somewhere between the previous two types and will arrive at the other side of the carnival at intermediate points in time. A snapshot of the carnival an hour after the arrival of the group would likely show some people already near the other side, others in the middle, and some still very close to the entrance.

Chromatography works in a similar way to the carnival analogy. A mixture known as the **analyte** (the initial group in our analogy) is subjected to certain conditions under which each component of the analyte behaves differently. These behavioral differences are the result of variation in size, polarity, charge, etc, and allow the components to be separated. There exist many different types of chromatography—paper, thin-layer chromatography, column, gas, etc.—that are used depending on the analyte and goals of the separation.

However, all types of chromatography are united by the existence of a **stationary phase** and a **mobile phase**. The analyte is dissolved in a fluid called the mobile phase that carries the analyte through a structure in which it will interact with a fixed substance known as the stationary phase (the vendors and rides in our analogy). Substances dissolve into the mobile phase to different degrees depending on their solubilities, while each component of the analyte also interacts differently with the stationary phase. The combination of these two factors allows many molecules to be separated due to each substance moving through the stationary phase at a different rate.



Yew Needle Paper Chromatography, Stefan Walkowski, CC BY-SA 4.0

The mixture of pigments in the needles of a Yew tree was the analyte for this chromatogram. Individual pigments were separated out using a 10% acetone in gasoline solution as the mobile phase that carried each pigment through the paper. Of the visible pigments, chlorophyll b made the least progress and is thus the lowest green pigment. Xanthophylls made the most progress and are the yellowish pigments towards the top.

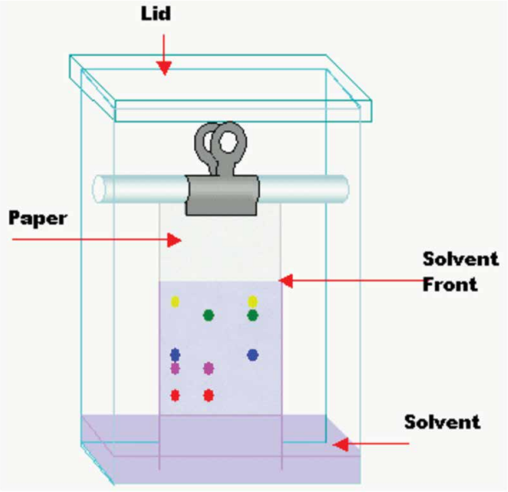
OBJECTIVE 7

Demonstrate an understanding of how chromatography can be used to separate the components of mixtures by taking advantage of variation in intermolecular forces between molecules.

Laboratories and advanced chemistry courses often use gas chromatography and **High-Performance Liquid Chromatography (HPLC)** to separate substances. In gas chromatography, the analyte is vaporized (turned into a gas) and put into a mobile phase of an unreactive gas like helium or argon. The mobile phase is then passed through a column containing solids that themselves act as the stationary phase or whose surfaces are covered with a liquid that acts as such. HPLC uses a liquid mobile phase that is passed through a column containing solid stationary phase at very high pressures. Both of these column chromatography methods are generally superior in many ways. For practical reasons, however, the two types of chromatography most seen in introductory chemistry courses are **paper chromatography and thin-layer chromatography (TLC)**.

Paper Chromatography

In ascending paper chromatography, the analyte is placed relatively low (but not at the very bottom) on a strip of paper made of cellulose, which along with its associated water molecules serves as the stationary phase. This process is often called "spotting" and is done by depositing the analyte using a small, glass capillary tube. The strip of paper is then lowered into a solvent that will begin to move up the paper through capillary action. The choice of solvent depends on the analyte, but water, alcohols, acetic acid, and acetone are common choices. It is important, however, that the solvent does not chemically react with components of the analyte.



Chromatography, Theresa Knott CC BY-SA 3.0

Once the **solvent front** (leading edge of the solvent) reaches the analyte, its components will dissolve in the solvent to varying degrees and be carried through the stationary phase. As the solvent continues moving upwards, the components of the analyte will move at different rates through the stationary phase, eventually creating a **chromatogram** in which each component is found at a different location on the paper. Polar substances adhere significantly to the stationary phase in paper chromatography. Such adhesion slows down their upward progress.

OBJECTIVE 7

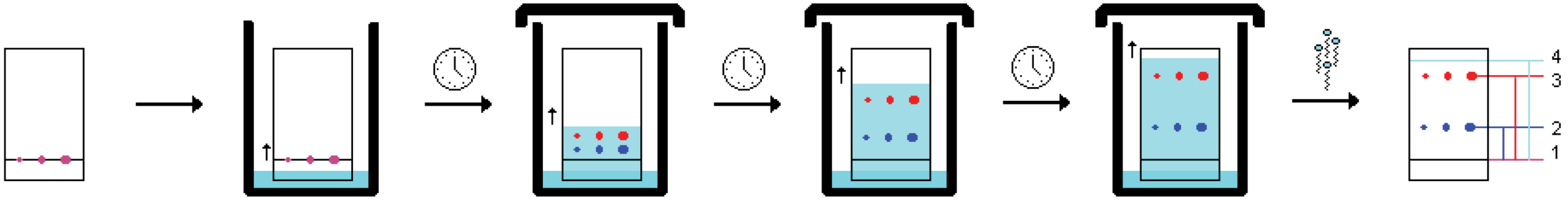
Demonstrate an understanding of how chromatography can be used to separate the components of mixtures by taking advantage of variation in intermolecular forces between molecules.

Thin-layer Chromatography (TLC)

Normal phase thin-layer chromatography works in a similar fashion to paper chromatography. The key difference is that the stationary phase is a thin layer of a substance like silica (SiO2) or alumina (Al2O3) that has been mounted to a strip of plastic or glass. TLC is commonly used due to its ease, low cost, relative quickness, and reliable separation for many substances. The silica and alumina typically used for the stationary phase for TLC are very polar and capable of hydrogen bonding. **Substances that are more polar will interact with the silica or alumina stationary phase more significantly and thus move up the plate the least.**

The choice of mobile phase depends on the polarity of the substances you wish to separate. Nonpolar solvents will carry nonpolar molecules through the stationary phase while often leaving polar molecules behind. Strongly polar solvents often just move nonpolar molecules to the solvent front with minimal separation while allowing for separation of polar molecules. Ideally, the solvent will be able to move all molecules without pushing them along the solvent front to maximize separation. The most common solvent is a mixture of ethyl acetate and hexane. This mixture of a moderately polar and a very nonpolar substance generally allows for good separation of many substances. However, there are many different solvent combinations that can be used depending on the circumstances.

One thing to note is that it is very common for the mixture you wish to separate to be colorless or at least have some colorless components. In this case, visualization has to be enhanced by some means. One of the approaches is the use of UV-light to detect certain substances on the TLC plate. In this case, the substances are not chemically altered. Another approach is the use of stains to identify particular substances on the TLC plate. Chemicals that react with the separated components of the analyte can give them color that allows them to be clearly visible. In this case, the substances are chemically altered.

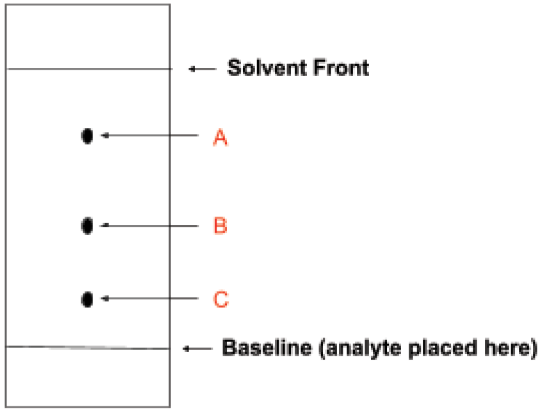


TLC Sequence, Courtesy of Wikimedia

OBJECTIVE 7

Demonstrate an understanding of how chromatography can be used to separate the components of mixtures by taking advantage of variation in intermolecular forces between molecules.

TASK 13: Rank the substances on the TLC plate shown below by polarity. Remember that the silica of the plate is very polar and capable of hydrogen bonding. More polar substances should experience stronger intermolecular interactions with the silica on the plate. The mobile phase used was a 6:1 hexane:ethyl acetate mixture.



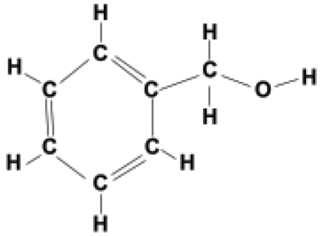
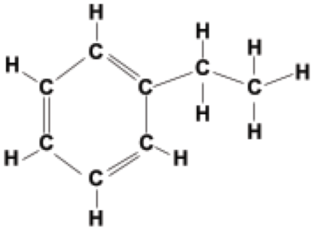
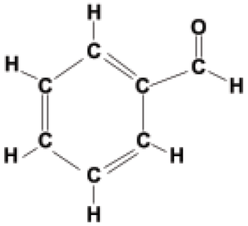
Most Polar	
Least Polar	

LOCK IT IN:
Briefly explain how you determined your ranking.

OBJECTIVE 7

Demonstrate an understanding of how chromatography can be used to separate the components of mixtures by taking advantage of variation in intermolecular forces between molecules.

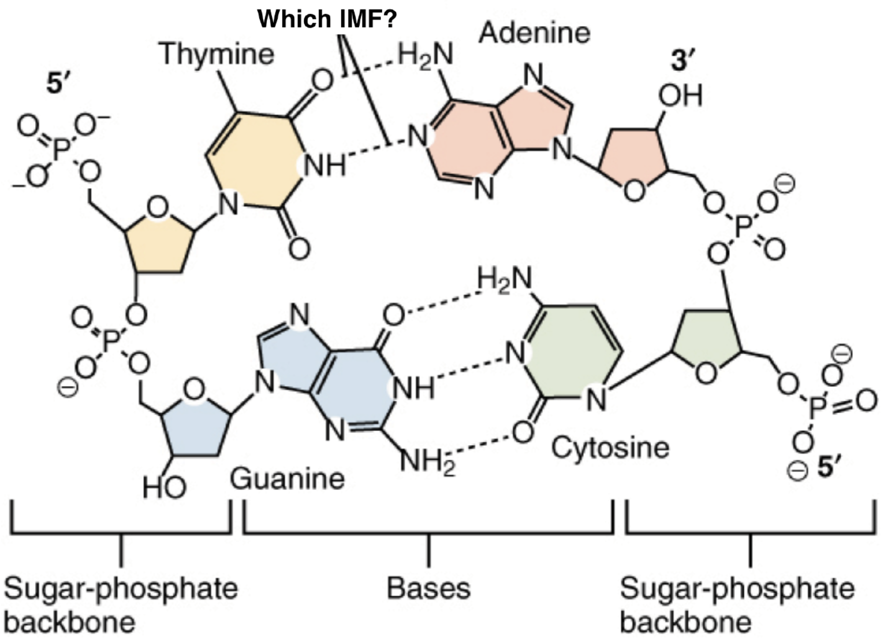
TASK 15: Use the structure of each substance below to match it to one of the lettered substances on the TLC plate from Task 13. Justify your answer choices in the appropriate space. **Hint:** Study the structure and composition of each molecule to determine the intermolecular forces it will have.

Substance	Matching Letter	Justification
 **Benzyl Alcohol**		
 **Ethylbenzene**		
 **Benzaldehyde**		

CLOSURE

CLOSURE: To help close the Intermolecular Forces Extension Activity and emphasize the importance of the concepts, let's apply your understanding of polarity and intermolecular forces to some concepts in biology. Answer the questions below using the associated images.

1. The image below shows two strands of the very famous molecule deoxyribonucleic acid, or DNA. A very important intermolecular force exists between the two strands and helps hold them together. Use your understanding of IMFs to identify the intermolecular force depicted in the image below of the two DNA strands. Explain how you know.



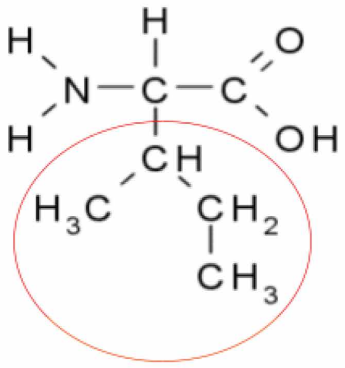
DNA Nucleotides, Open Stax College CC BY 3.0

CLOSURE

2. Would DNA be a water soluble molecule? Explain your answer.

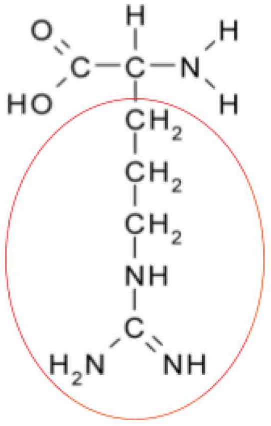
3. Protein folding is another application of intermolecular forces in biology. Dispersion forces, dipole-dipole interactions, hydrogen bonding, and even ion-dipole interactions all play key roles in determining how proteins fold into the extraordinary number of shapes that they do. Different amino acids are more suited to different locations in a protein based on the protein's location in an organism and its function. Consider a protein that will spend most of its time surrounded by water. Which of the amino acids below would be best suited for the exterior of such a protein in that it would interact most favorably with water molecules? Justify your reasoning. Hint: Focus on the circled side chains.

Isoleucine



CC BY-SA 3.0

Arginine



CC BY-SA 3.0