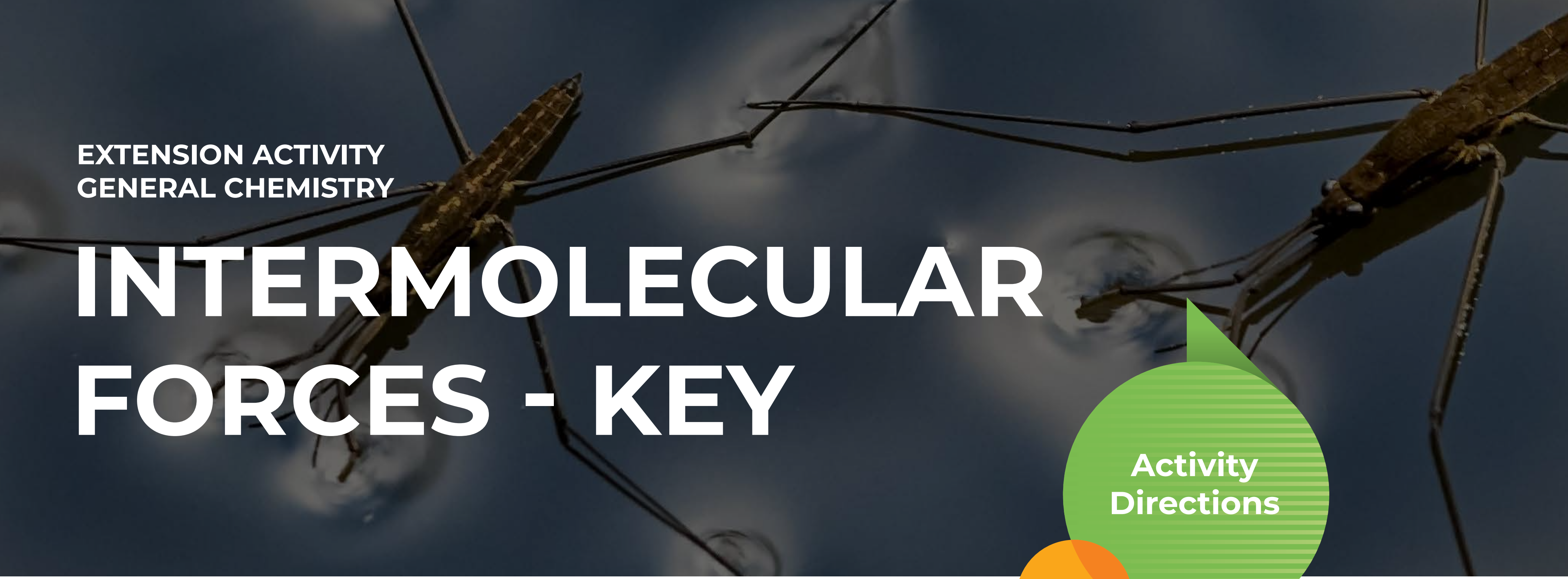


EXTENSION ACTIVITY
GENERAL CHEMISTRY

INTERMOLECULAR FORCES - KEY

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.

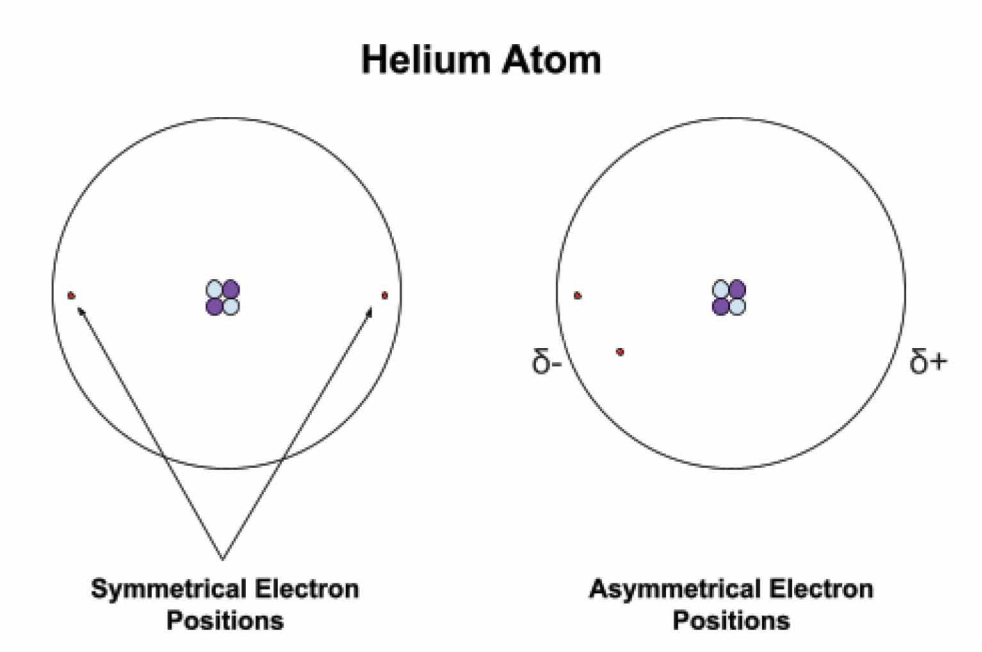
Activity Directions

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.

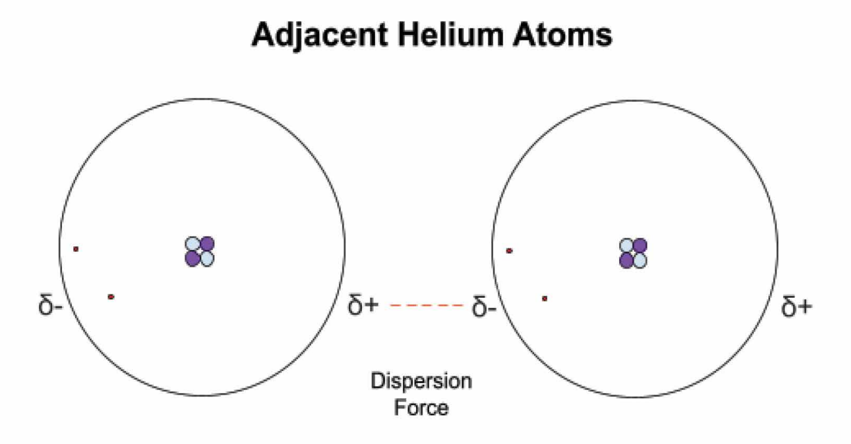
One important concept to remember in chemistry is that atoms, molecules, and their components are always in motion. This motion gives rise to very important phenomena, one of which is demonstrated by London dispersion forces. Although the Bohr molecule of the atom which shows electrons moving around 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**.


Adjacent Helium Atoms

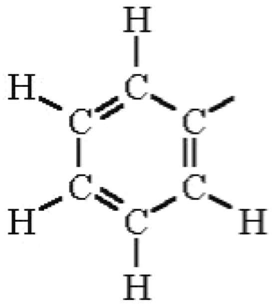
The dispersion force will be the strongest intermolecular force in substances that do not have permanent dipoles. Although dispersion forces can be quite substantial as molecules increase in their number of electrons, permanent dipoles generate stronger forces in many small molecules. An easy way to 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.

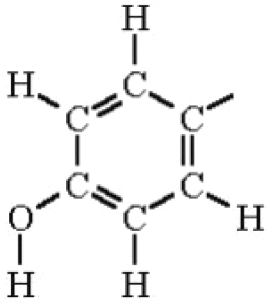


Phenylalanine Side Chain

Chemical Bond	Electronegativity of First Element	Electronegativity of Second Element	Electronegativity Difference	Polar or Nonpolar?
C-C or C=C	2.55	2.55	0	Nonpolar
C-H	2.55	2.20	0.35	Nonpolar

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	2.55	2.55	0	Nonpolar
C-H	2.55	2.20	0.35	Nonpolar
C-O	2.55	3.44	0.89	Polar
O-H	3.44	2.20	1.24	Polar

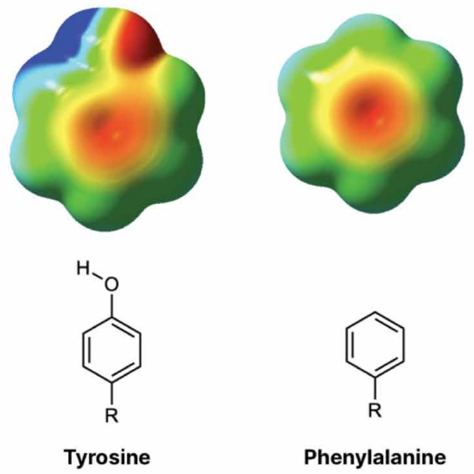
LOCK IT IN:
For which of the two side chains shown above will London Dispersion Forces be the only intermolecular force in its interactions? Explain your answer.

London dispersion forces will be the predominant intermolecular force for the phenylalanine side chain. The tyrosine side chain contains polar bonds between carbon and oxygen and oxygen and hydrogen. These bonds give tyrosine's side chains dipole-dipole interactions. Phenylalanine's side chain only contains nonpolar bonds and thus will interact through dispersion forces.

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

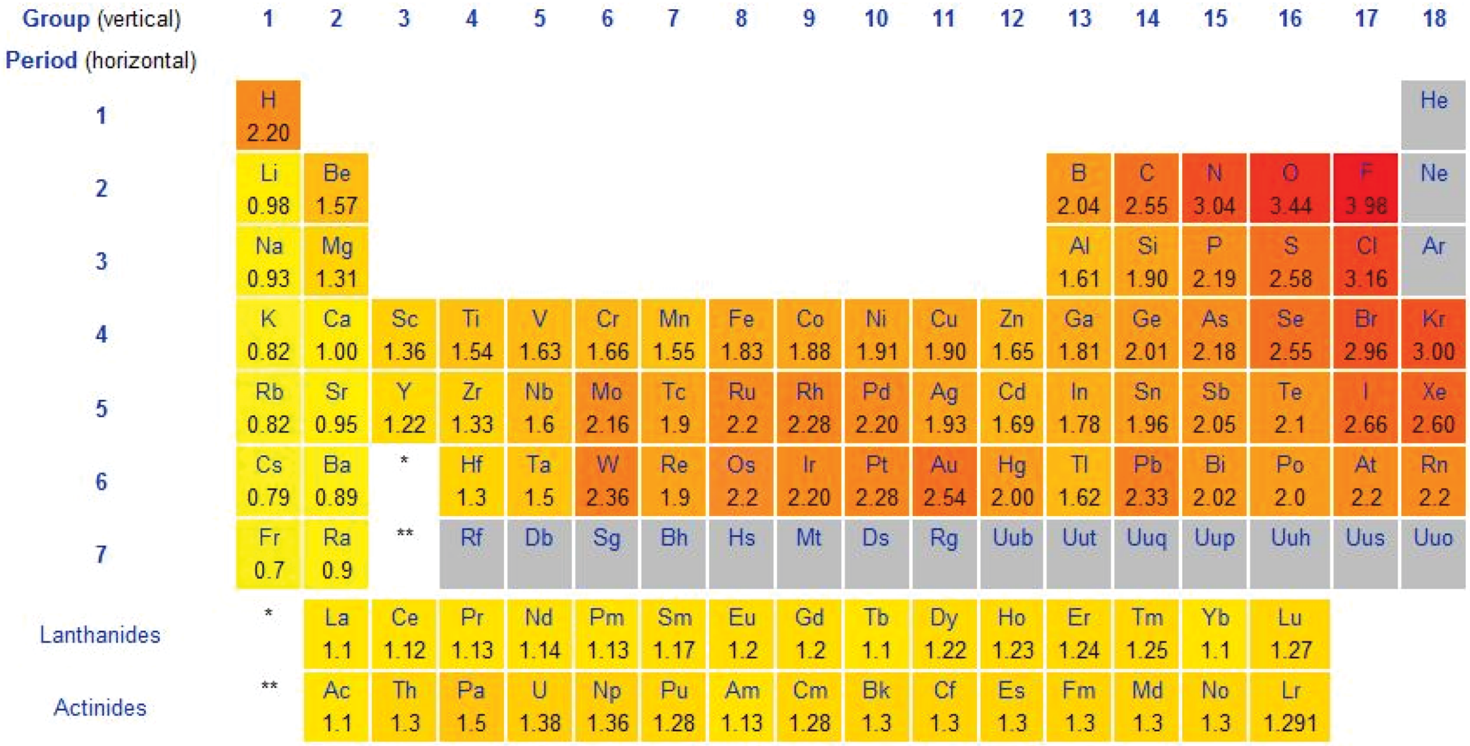
LOCK IT IN:
Do the electrostatic potential maps confirm your decision as to which side chain will have dispersion forces as its only intermolecular force? Why or why not?

The electrostatic potential maps demonstrate that the tyrosine side chain has polar regions that give it a net dipole whereas the phenylalanine side chain does not have those same regions.

OBJECTIVE 1

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

Pauling Electronegativities of the Elements



Group (vertical) Period (horizontal)	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
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.

The strongest intermolecular forces in the phosphorus sulfides will be London dispersion forces because all of the bonds will be nonpolar. The bonds are either between two of the same element or between phosphorus and sulfur. None of those bonds will lead to an electronegativity difference.

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.

Examples only. Answers will vary.

	Chemical Formula	Molar Mass (g/mol)	Monatomic, Homonuclear, or Compound?
Strongest LDF	CS_2	78	Compound
	Ar	40	Monatomic
	O_2	32	Homonuclear
	N_2	28	Homonuclear
	CH_4	16	Compound
Weakest LDF	He	4	Monatomic

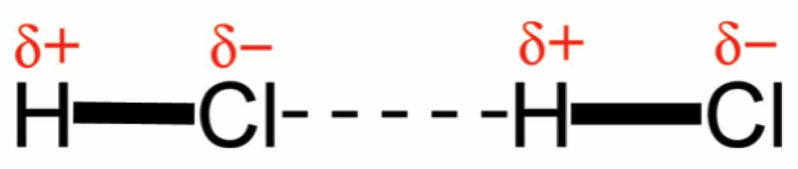
LOCK IT IN:
Explain what reasoning you used to create molecules only showing London dispersion forces and how you were able to rank the strength of those forces.

To make molecules only showing London dispersion forces one should either look for monatomic substances, homonuclear substances, or compounds in which the molecules contain only nonpolar bonds or dipoles that cancel out.

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.



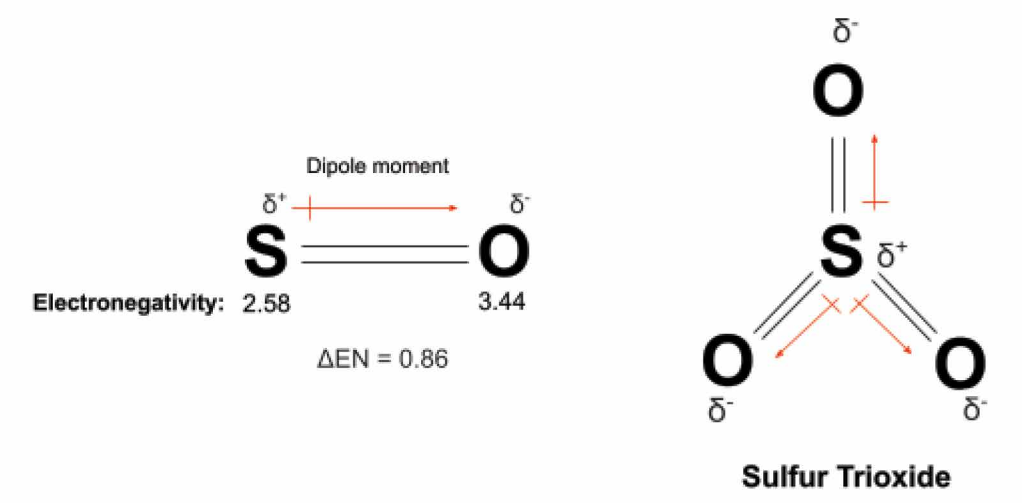
The permanent dipole in HCl due to electronegativity differences between hydrogen and chlorine creates the dipole-dipole interaction indicated by the dashed line in the image.

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.

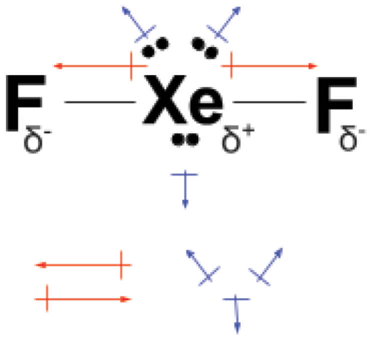
OBJECTIVE 2

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

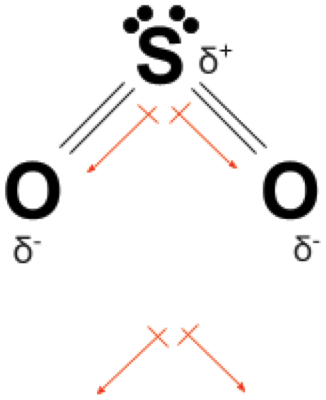


Dipoles Cancel
Each Other Out

Nonpolar Molecule

Sulfur Dioxide

Molecular Geometry: Bent



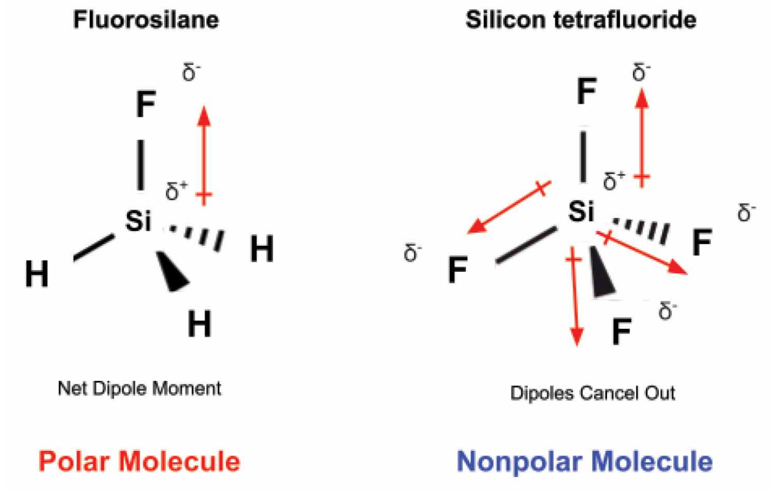
Dipoles Do Not Fully Cancel
Net Dipole Moment Present

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	2.55	3.04	0.49	Polar	Yes
S — O	2.58	3.44	0.86	Polar	Yes
Si — C	1.90	2.55	0.65	Polar	Yes
I — Br	2.66	2.96	0.30	Nonpolar	No
P — O	2.19	3.44	1.25	Polar	Yes

LOCK IT IN:
Chlorine monofluoride (ClF) 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.

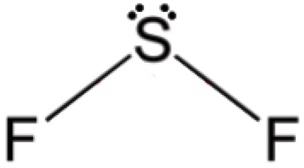
Since there are only two atoms in the molecule, determining if the bond is polar is enough to determine that there is a net dipole.* The difference in electronegativity between chlorine and fluorine is 0.82—sufficient for classification as a polar bond and sufficient for classification of chlorine monofluoride as a polar molecule.

***Interestingly enough, carbon monoxide only contains two atoms, but the polarity of its triple bond is not accurately predicted by electronegativity differences. Although it does still have a net dipole, the carbon atom has a greater electron density than the oxygen atom. This phenomenon is due to reasons you should research if you are curious.**

OBJECTIVE 2

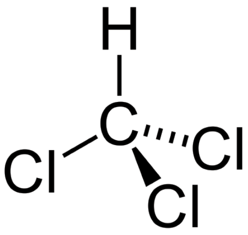
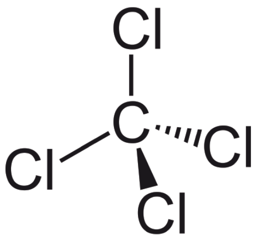
Demonstrate an understanding of how dipole-dipole interactions form and vary amongst different molecules.

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≡N Linear	Yes	Nitrogen is significantly more electronegative than carbon and thus that bond has a dipole moment towards nitrogen. However, this dipole moment is not canceled out by the carbon—hydrogen bond which is nonpolar.
Carbon Dioxide	CO_2	O=C=O Linear	No	Although the carbon—oxygen bond is very polar, the symmetry of the linear molecule prevents a net dipole from existing for the molecule.
Sulfur Difluoride	SF_2	 Bent	Yes	The sulfur—fluorine bond is very polar due to fluorine's high electronegativity compared to that of sulfur. However, the bent symmetry of the molecule does not cancel out the dipoles of each bond and thus make sulfur difluoride polar.

OBJECTIVE 2

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

Chloroform	$CHCl_3$	 Tetrahedral	Yes	The high electronegativity of chlorine compared to carbon makes each of the bonds connecting them polar. The presence of one hydrogen atom in chloroform disrupts the symmetry of the molecule and allows for the formation of a net dipole.
Carbon Tetrachloride	CCl_4	 Tetrahedral	No	Although the bonds between chlorine and carbon are polar, the symmetry of the tetrahedral molecular geometry cancels out the dipoles. As a result, there is no net dipole for carbon tetrachloride.

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.

Even if the bonds in a molecule are polar, the dipole moments of these bonds can cancel each other out if the molecule is symmetrical. In the case of carbon tetrachloride, the bonds are polar. The tetrahedral symmetry, however, means that the four bond dipole moments cancel each other out leaving no net dipole.

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

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

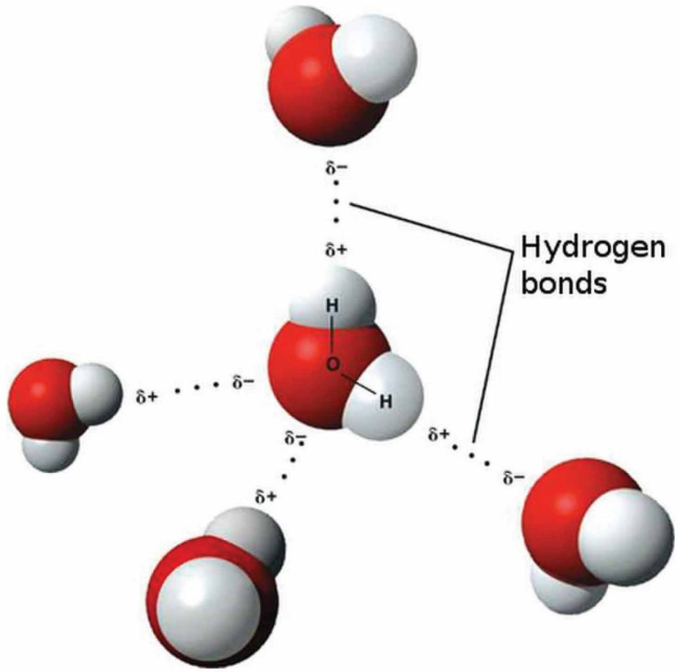
Fluorine, oxygen, and nitrogen form the most polar bonds with hydrogen.

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.



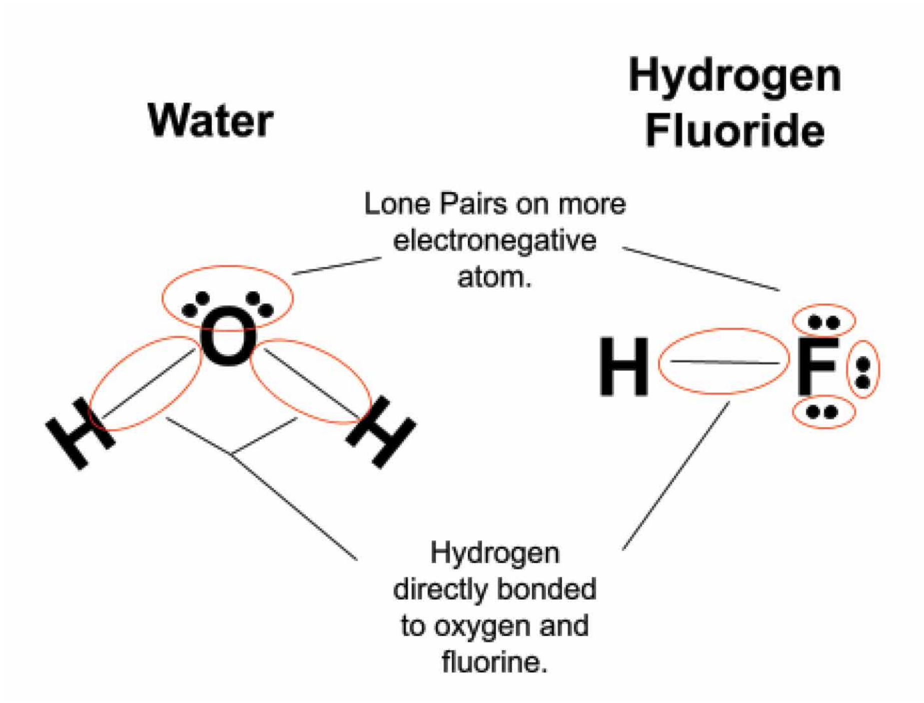
Hydrogen bonds depicted between water molecules.

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.

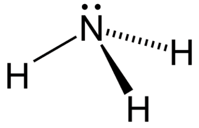
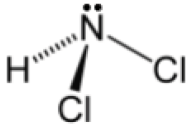
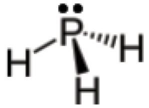
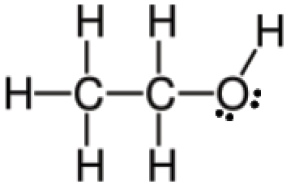
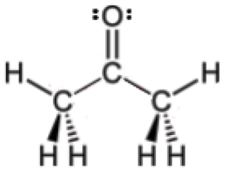
Look at the structures of water and hydrogen fluoride, both of which are capable of hydrogen bonding:



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		Yes
Dichloramine	$NHCl_2$		Yes
Phosphine	PH_3		No
Ethanol	C_2H_6O		Yes
Acetone	C_3H_6O		No

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.

You should look for bonds DIRECTLY between hydrogen and either oxygen, fluorine, or nitrogen. You should then look for lone pairs on the more electronegative atom.

2. Comparing Hydrogen Bonding Capabilities—Polarity

Since hydrogen bonding plays an important role in the properties of many molecules, it is sometimes helpful to compare hydrogen bonding capabilities. One factor that affects a molecule's ability to participate in hydrogen bonding is the polarity of the relevant bonds. Hydrogen bonded to nitrogen will be less partially positive than when it is bonded to fluorine. This is because nitrogen is less electronegative than fluorine. Thus the ranking of bond dipole moments is as follows:

F — H > O — H > N — H

One important effect of hydrogen bonding in molecules is that they typically increase the boiling point of substances in which they occur. This concept will be explored in much more detail in the Latent Heat Game, but we will use it a little here. If one simply considers bond polarity, hydrogen fluoride should have a higher boiling point than water given that they have similar molar masses. Let's see if that is the case.

Molecule	Molar Mass (g/mol)	Standard Boiling Point (°C)
Water (H_2O)	18.0	100.0
Hydrogen Fluoride (HF)	20.0	19.5
Hydrogen Sulfide (H_2S)	34.1	— 60.2
Hydrogen Chloride (HCl)	36.5	— 85.1

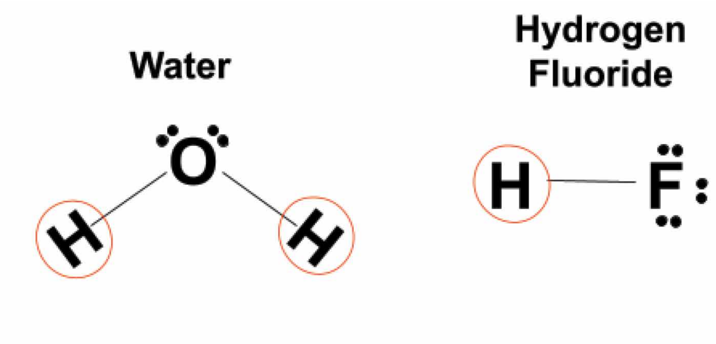
Please note that both substances have much higher boiling points than similar molecules as a result of hydrogen bonding. Hydrogen sulfide is analogous to water and has greater dispersion forces due to its larger size, but its boiling point is much lower. The same is true for hydrogen chloride which is analogous to hydrogen fluoride. However, water's boiling point is significantly higher than that of hydrogen fluoride despite the bond between hydrogen and fluorine being more polar than the one between hydrogen and oxygen. **Although the polarity of the bond is important, there is a more important factor at work.**

OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.

3. Comparing Hydrogen Bonding Capabilities—Extensiveness

One of the most important factors in comparing the hydrogen bonding capabilities of molecules is a concept that is sometimes called **extensiveness**. Whereas bond polarity focuses on the magnitude of the dipole, extensiveness is merely a description of *how many* hydrogen bonds a molecule is capable of forming. Extensiveness plays a larger role in hydrogen bonding than polarity and requires a few steps to assess:

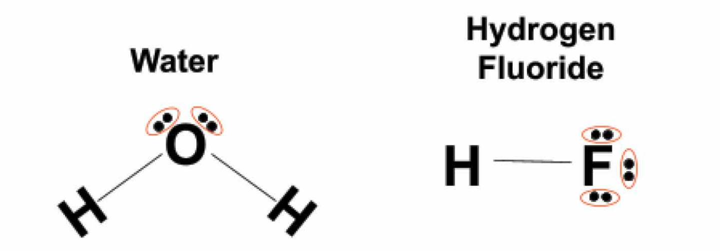


Determine how many hydrogen atoms will participate in hydrogen bonding.

Simply count the number of hydrogen atoms that are directly bonded to the electronegative atoms in the molecule. These will form the partial positive ends of hydrogen bonds with other molecules. Notice that a single water molecule is able to provide two hydrogen atoms for bonding, while a single hydrogen fluoride molecule is only able to provide one.

Determine how many lone pairs are available to participate in hydrogen bonding.

Each lone pair on the highly electronegative atom with the partial negative charge is able to contribute to hydrogen bonding. Notice that the oxygen atom in water has two lone pairs of electrons. Each lone pair can participate in hydrogen bonding with a separate molecule. Thus the one oxygen atom in water can actually contribute to two hydrogen bonds. The fluorine in hydrogen fluoride, however, has three lone pairs that can participate in hydrogen bonding.



OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.

Determine the average number of hydrogen bonds the molecule can form.

If we were to simply sum the number of hydrogen bonds that individual molecules of water and hydrogen fluoride could make, we would get four for both. Water can provide two hydrogen atoms and has two lone pairs on the oxygen that can contribute to hydrogen bonding as well. Hydrogen fluoride can provide one hydrogen atom and has three lone pairs on the fluorine that can contribute to hydrogen bonding.

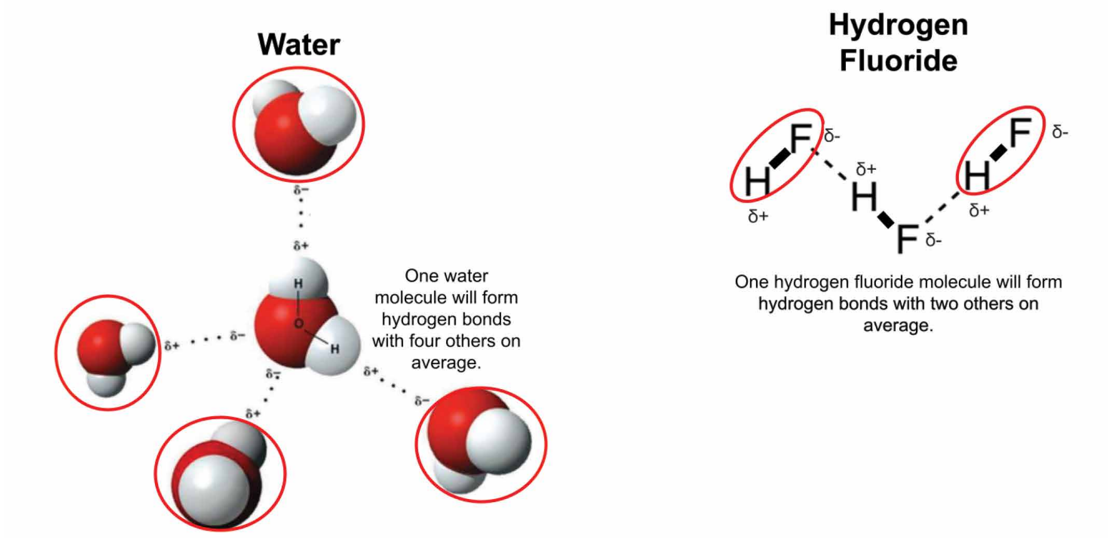
However, the hydrogen bonding capabilities of these molecules are not equal. Despite having three lone pairs capable of hydrogen bonding, the fluorine atom in hydrogen fluoride will only contribute to one hydrogen bond on average. This is because other hydrogen fluoride molecules only have one partially positive hydrogen to contribute. The usefulness of the lone pairs is limited by the number of potential hydrogen partners. As a result, a single hydrogen fluoride molecule will only form two hydrogen bonds on average with other hydrogen fluoride molecules.

Water, however, demonstrates a balance of sorts. Each water molecule can contribute the partial positive portion of the hydrogen bond to two other water molecules, while the two lone pairs can contribute the partial negative portion to two other molecules. As a result, a single water molecule forms four hydrogen bonds on average with other water molecules—twice as many as hydrogen fluoride. *It is this difference that explains why water's boiling point is significantly higher than that of hydrogen fluoride.*

Water and hydrogen fluoride demonstrate that when comparing hydrogen bonding capabilities, it is important to remember that more lone pairs does not necessarily mean more hydrogen bonds on average. Lone pairs will only be as useful to hydrogen bonding as there are hydrogen partners with which to interact. The same is true in reverse. More hydrogen atoms available for bonding will only be as useful as there are lone pairs with which they can interact.

OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.



One water molecule will form hydrogen bonds with four others on average.

One hydrogen fluoride molecule will form hydrogen bonds with two others on average.

If the number of relevant hydrogen atoms and number of relevant lone pairs is equal:

Average # of Hydrogen Bonding Partners = Number of Lone Pairs + Number of Partially Positive Hydrogen Atoms

Water HB Partners = 2 Lone Pairs + 2 Hydrogen Atoms

Water HB Partners = 4

If the number of relevant hydrogen atoms and number of relevant lone pairs is different:

Average # of Hydrogen Bonding Partners = 2 x Smaller Number (lone pairs OR partially positive hydrogens)

Hydrogen Fluoride HB Partners = 2 x 1 partially positive hydrogen

Hydrogen Fluoride HB Partners = 2

OBJECTIVE 3

Demonstrate an understanding of why hydrogen bonding occurs.

It is also very important to note that hydrogen bonding occurs in addition to other intermolecular forces. Although hydrogen bonding helps contribute to many patterns, particularly those covered in the Latent Heat Game, they do not necessarily surpass the contributions of other forces and interactions. For example, a small molecule capable of hydrogen bonding will likely still have a lower boiling point than a really large molecule with only London dispersion forces. As such, IMF comparisons are easiest amongst molecules of similar sizes and become more complex as the difference in size and composition between molecules increases.

TASK 7: Compare the hydrogen bonding capabilities of the molecules you identified as hydrogen bonding capable from Task 6 to the hydrogen bonding capabilities of water. Simply write in the name of the molecules into the column on the right and use a ">" or "<" symbol to indicate which molecule has the greater hydrogen bonding capabilities. Although there are five spaces, you do not necessarily need to fill all spaces. Only use the molecules you identified as capable of hydrogen bonding. Remember to use both your understanding of extensiveness and polarity to make your decision!

Substance #1	Relative Hydrogen Bonding Capability (< or >)	Substance #2
Water (H_2O)	>	Ammonia (NH_3)
	>	Dichloramine ($NHCl_2$)
	>	Ethanol (C_2H_6O)

LOCK IT IN:
Generally describe how water's hydrogen bonding capabilities compare to the other molecules provided.

Water has greater hydrogen bonding capabilities than all of the molecules provided.

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.

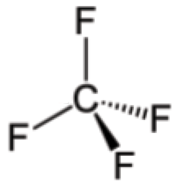
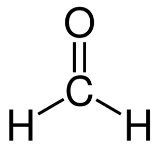
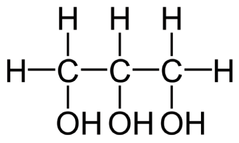
You have certainly noticed that oil and water do not mix. This is the result of them being insoluble in one another. A term more commonly used to describe the two is **immiscible. Miscibility** describes the ability of two substances to mix in all ratios due to their high solubilities in one another. 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 8: 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		London dispersion forces	insoluble
Formaldehyde	CH_2O		Dipole-Dipole	soluble
Neon	Ne	Ne	London dispersion forces	insoluble
Dioxygen	O_2	$O=O$	London dispersion forces	insoluble
Glycerol	$C_3H_8O_3$	 $H-\underset{OH}{\overset{H}{C}}-\underset{OH}{\overset{H}{C}}-\underset{OH}{\overset{H}{C}}-H$	Hydrogen Bonding	miscible

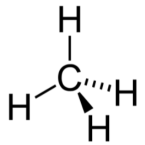
OBJECTIVE 4

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

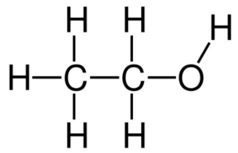
LOCK IT IN:
Take note of the substances in the table that are most and least soluble in water. What types of intermolecular forces seem to increase the solubility of a substance in water? Why does that make sense?

Dipole-dipole interactions and hydrogen bonding increase the solubility of a substance in water. This is because water is a polar molecule capable of hydrogen bonding. Since "like dissolves like", it makes sense that other polar and hydrogen bonding-capable molecules will often be soluble in water.

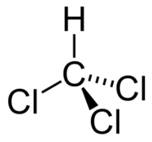
TASK 9: In the box to the right are three substances that you have encountered as completing this extension activity. Use their structures to rank them in terms of their solubility in water and then provide a brief justification as to your ranking.



Methane
CH_4



Ethanol
C_2H_6O



Chloroform
$CHCl_3$

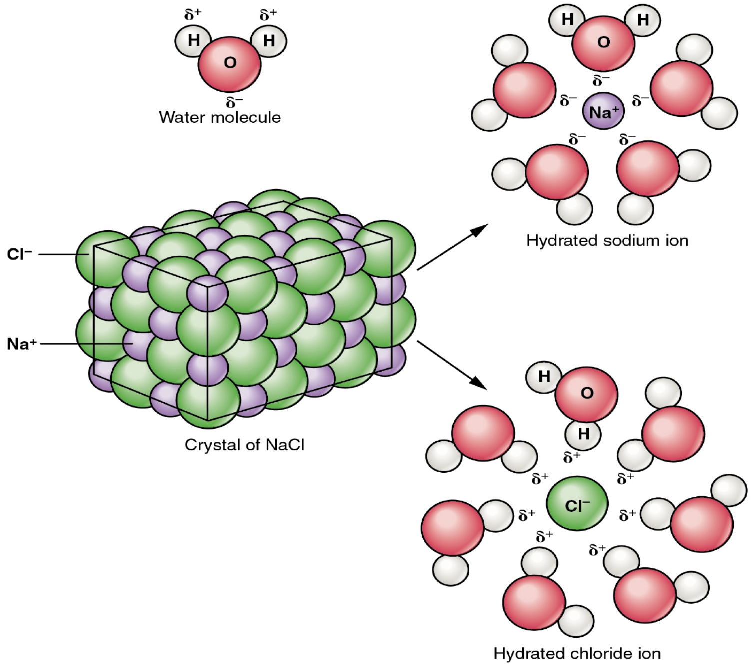
	Chemical Name and Formula	Justification
Most Soluble in Water	Ethanol	Ethanol is capable of forming hydrogen bonds with water, which increases its solubility to the point that it is miscible.
	Chloroform	Chloroform is polar, which increases its solubility in water though not as much as ethanol.
Least Soluble in Water	Methane	Methane is nonpolar and incapable of hydrogen bonding. Its solubility in water is very low.

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

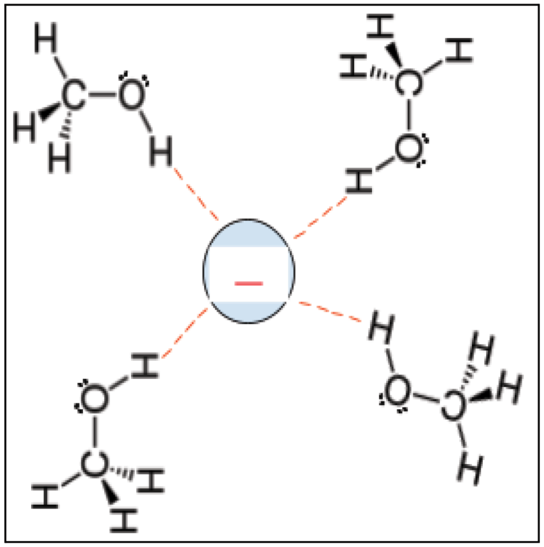
The lattice energy of ionic compounds is not easily overcome, but the dipoles of polar solvents are often capable of doing so because of how they arrange themselves around individual ions to create energetically favorable conditions within a solution. For example, the polarity of water molecules are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

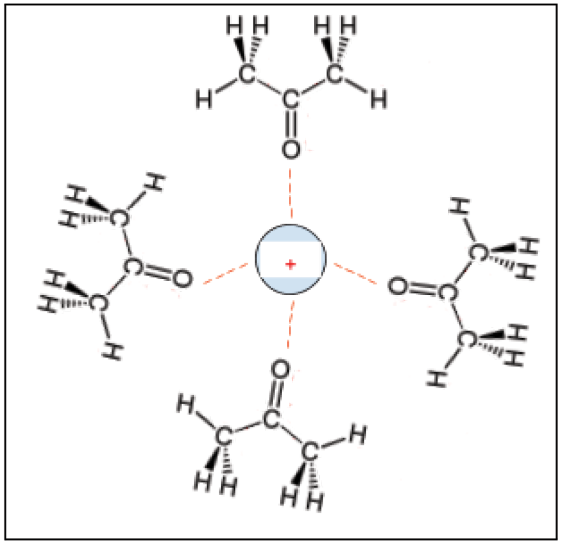
The image to the left depicts water molecules forming hydration shells around sodium and chloride ions that have been separated from one another. Notice how the oxygen atom bearing the partial negative charge is oriented towards the positively charged sodium ion, while the hydrogen atoms with the partial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

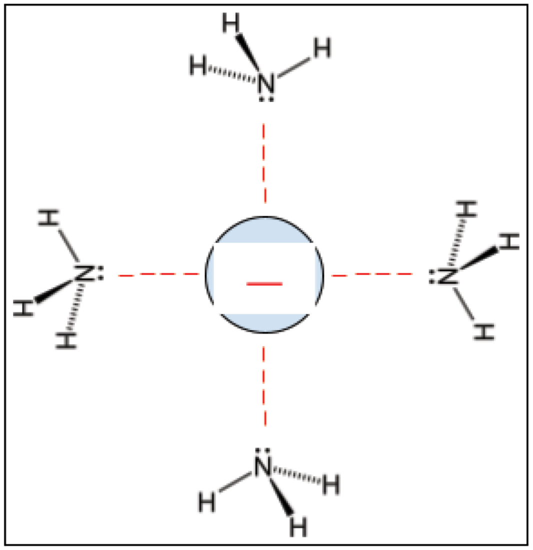
OBJECTIVE 5

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

TASK 10: 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.



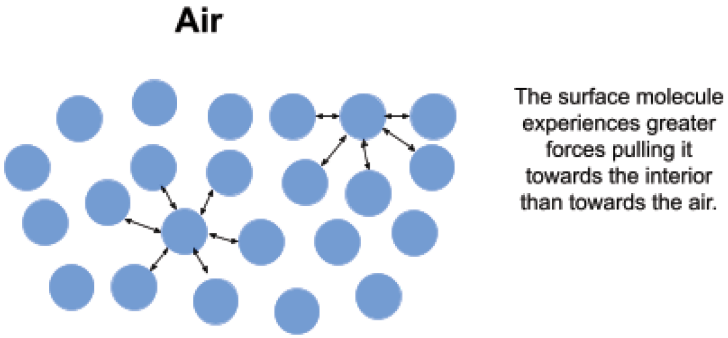




OBJECTIVE 6

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.



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.

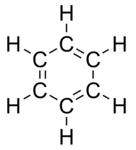
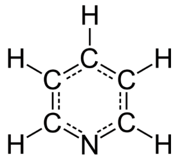
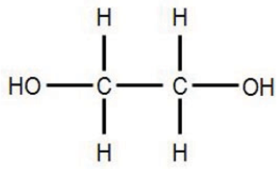
OBJECTIVE 6

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 11: 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		0.028	Benzene is nonpolar and has more modest intermolecular forces pulling molecules toward the interior of the liquid. It has the lowest surface tension as a result.
Pyridine	C_5H_5N		0.037	Pyridine is polar due to the nitrogen atom in its ring that gives it dipole-dipole interactions and increases its surface tension compared to benzene.
Ethylene Glycol	$C_2H_6O_2$		0.048	The presence of the hydroxyl groups (—OH) makes ethylene glycol not only polar, but also capable of hydrogen bonding. These interactions lead to a higher surface tension than those of pyridine and benzene.

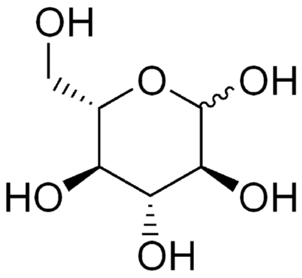
OBJECTIVE 6

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

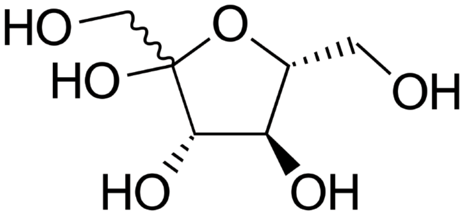
A key characteristic of liquids is their ability to flow. However, liquids like water will flow much more easily than liquids like honey. The term used to describe a liquid's resistance to flow is **viscosity**. As such, we say that honey is more viscous (has a higher viscosity) than water. Viscosity 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



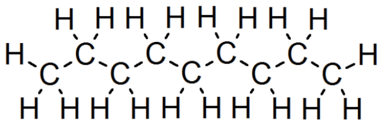
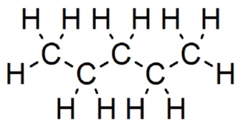
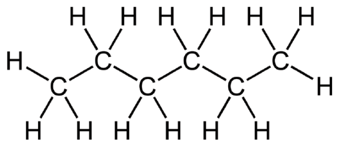
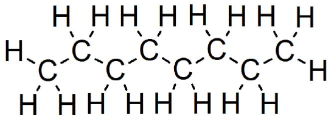
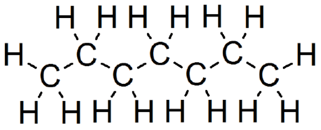
Hydrogen bonding between the sugar molecules of glucose and fructose help give honey its high viscosity.

OBJECTIVE 6

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

TASK 12: 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	0.665
n-Pentane	C_5H_{12}		72.15	0.224
n-Hexane	C_6H_{14}		86.17	0.295
n-Octane	C_8H_{18}		114.2	0.509
n-Heptane	C_7H_{16}		100.2	0.389

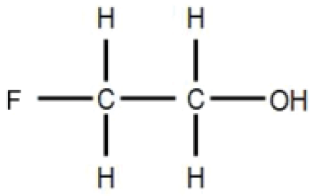
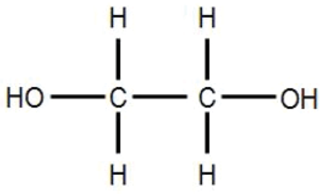
OBJECTIVE 6

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.

Since all of these molecules have similar geometry and contain only nonpolar bonds, comparing their viscosities only requires comparing the strength of their London dispersion forces. Since London dispersion forces generally increase with the number of electrons in the molecule and all of these molecules are composed of carbon and hydrogen, we can give the highest viscosity to the largest molecule (n-nonane) and decrease from there.

TASK 13: Take a look at the two molecules with very similar molar masses below. In the space beside the table, explain which has the greater viscosity using your knowledge of intermolecular forces.

		
Chemical Name	2-Fluoroethanol	Ethylene Glycol
Chemical Formula	C_2H_5FO	$C_2H_6O_2$
Molar Mass (g/mol)	64.1	62.1

Ethylene glycol would have a higher viscosity due to its ability to form more hydrogen bonds than 2-fluoroethanol. Ethylene glycol has both two partially positive hydrogens and two lone pairs on each of the oxygen atoms. 2-fluoroethanol does have one hydroxyl (—OH) group for hydrogen bonding, but the fluorine on the other side of the molecule can only serve as a hydrogen bond donor since it does not have a hydrogen to serve as an acceptor.

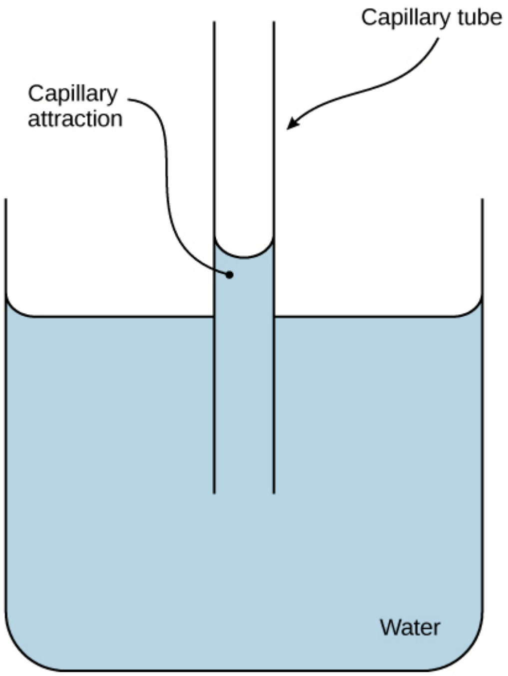
OBJECTIVE 6

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

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.



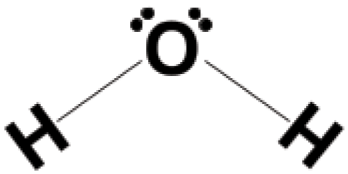
Capillary Action, Open Stax College CC BY 4.0

OBJECTIVE 6

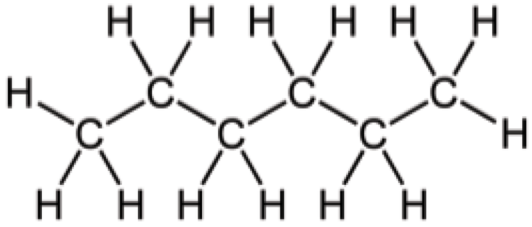
Demonstrate an understanding of how intermolecular forces influence surface tension, viscosity, and capillary action.

TASK 14: 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



The water will move up the glass tube higher than n-hexane due to the fact that the glass is polar which will allow for more adhesion with the polar water molecules. Hexane, however, is nonpolar and will not experience as significant of adhesive forces with the glass.

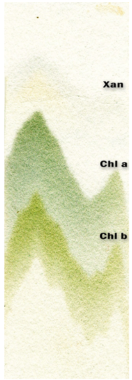
OBJECTIVE 7

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.

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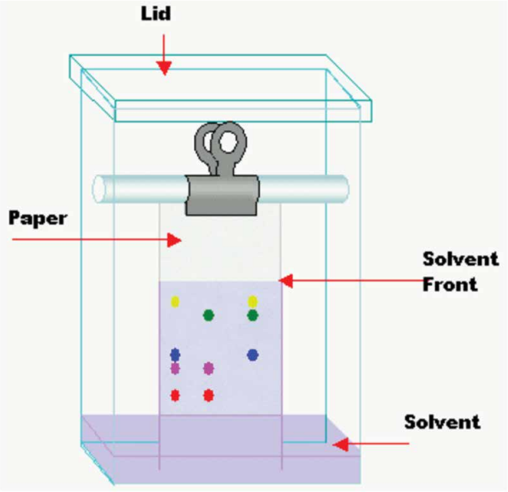
Laboratories and advanced chemistry courses often use gas chromatography and **High-Performance Liquid Chromatography (HPLC)** to separate substances. In gas chromatography, the analyte is vaporized (turned into a gas) and put into a mobile phase of an unreactive gas like helium or argon. The mobile 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.

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.



Chromatography, Theresa Knott CC BY-SA 3.0

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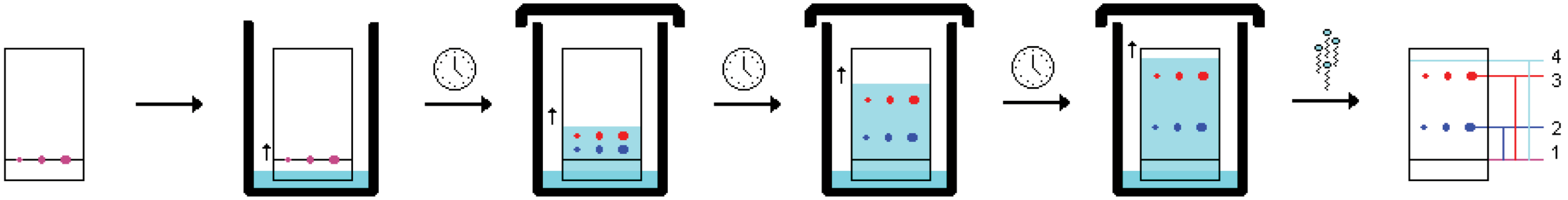
Thin-layer Chromatography (TLC)

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

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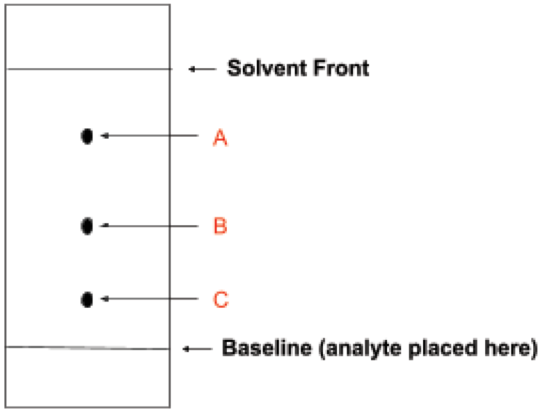
TASK 15: Rank the substances on the TLC plate shown below by polarity. Remember that the silica of the plate is very polar and capable of hydrogen bonding. More polar substances should experience stronger intermolecular interactions with the silica on the plate. The mobile phase used was a 6:1 hexane:ethyl acetate mixture.



Most Polar	C
	B
Least Polar	A

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

The most polar substance interacts the most with the polar silica of the stationary phase. As a result, it makes the least progress up the plate. Substance C has clearly moved the smallest distance upwards, indicating that it is the most polar. The opposite is true of substance A, which has moved up the plate the most.

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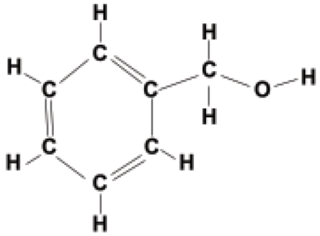
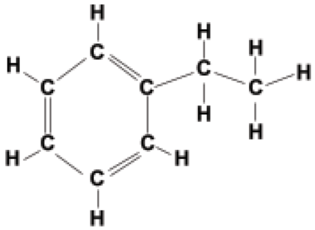
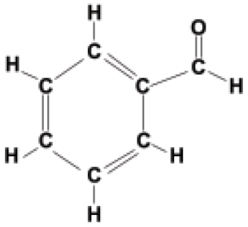
TASK 16: Use the structure of each substance below to match it to one of the lettered substances on the TLC plate from Task 15. Justify your answer choices in the appropriate space. **Hint:** Study the structure and composition of each molecule to determine the intermolecular forces it will have.

Substance	Matching Letter	Justification
 **Benzyl Alcohol**	C	Benzyl alcohol contains the polar carbon—oxygen and oxygen—hydrogen bonds, which increases its interactions with the polar TLC plate. The hydroxyl group (—OH) also allows for benzyl alcohol to form hydrogen bonds with the plate. The strong intermolecular interactions slow down the upward movement of the substance on the plate. As such, C is the correct choice.
 **Ethylbenzene**	A	Ethylbenzene is a nonpolar molecule and thus will interact the least with the polar silica of the TLC plate. It most quickly moves up the plate, as shown by substance A.
 **Benzaldehyde**	B	While benzaldehyde is certainly more polar than ethylbenzene, it is not able to form hydrogen bonds with the TLC plate due to hydrogen being bonded to the carbon instead of the oxygen. It moves up the plate at a rate intermediate that of ethylbenzene and benzyl alcohol. As such, B is the correct choice.

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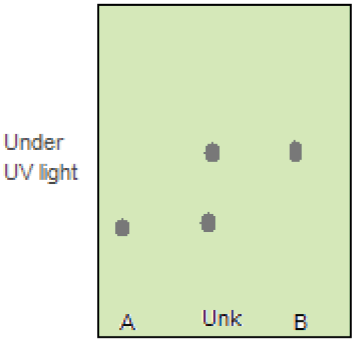
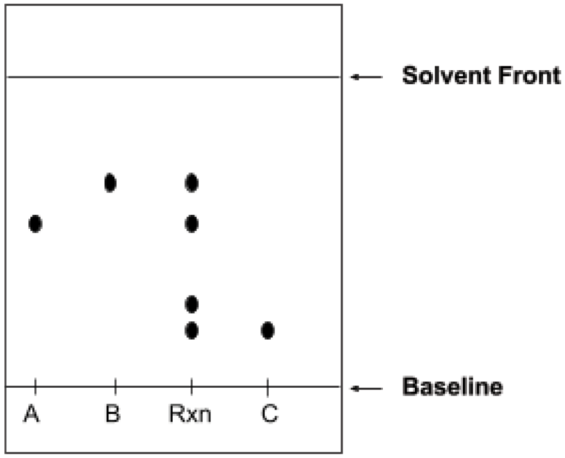
One way in which TLC can be useful is when one wants to synthesize a compound in the lab. TLC can help by confirming the presence of substance in a mixture through comparison with standard samples of that sample. Take a look at the TLC plate below that shows a standard sample of substance A run in the first lane, a standard sample of substance B in the third lane, and the mixture of interest in the middle lane labeled “Unk” for unknown. Notice that the analyte in the middle lane has spots equivalent to those of standard samples of substance A and B. The mixture can thus be said to contain both substance A and B.



TLC Results, CS8hu CC BY-SA 3.0

TASK 17: A chemist completes a synthesis reaction where A + B → C. After completing the reaction, they perform TLC with the results of their reaction and standard samples of substances A, B, and C for comparison. Study the TLC plate and answer the accompanying questions.



Was the scientist able to create the desired product? If not, explain how you know. If yes, were all of the reactant molecules (A and B) successfully converted to C? Do you notice anything else on the TLC plate that should be noted?

The scientist was able to create the desired product (substance C), as shown by the spot in the reaction lane that matches the standard sample of substance C. However, there are still spots that indicate the presence of the reactants A and B. This tells us that the reaction did not fully convert the reactants to the product. There is also a spot that does not match those of substance A, B or C, which indicates that another product was generated as well.

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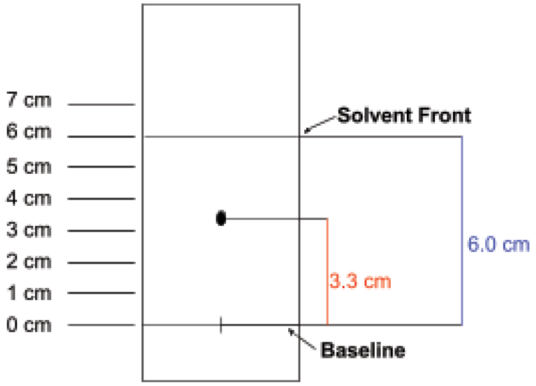
TLC can also be used to identify and compare substances quantitatively. After performing TLC, a chemist will often calculate the retardation factor or retention factor (Rf) for each spot on the plate. To calculate the Rf, the following formula is used:

$$R_f = \frac{\text{Distance Traveled By Sample}}{\text{Distance Traveled By Solvent}}$$

Using a ruler, one can measure the distance from where the analyte was placed (the baseline) to the center of the spot that has migrated upwards. This number will then be divided by the distance from the baseline to the solvent front. In normal phase TLC, the more polar substance will have the lower Rf because it will have traveled a shorter distance. A less polar substance will move upwards more and have a larger Rf. The value of Rf will always be a number from zero to one and dimensionless (no unit).



$$R_f = \frac{\text{Distance Traveled By Sample}}{\text{Distance Traveled By Solvent}} = \frac{3.3\text{ cm}}{6.0\text{ cm}} = \mathbf{0.55}$$

The retention factor of a substance is only consistent when the TLC conditions are consistent. Substances will have different Rf factors in different mobile phases or when interacting with different stationary phases. However, if TLC conditions are the same, a substance can often be identified by comparing its Rf with substances on a table of known retention factors for the stationary and mobile phases you used.

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TASK 18: A biochemist believes that they successfully synthesized two amino acids. They complete TLC with a stationary phase of silica and a mobile phase of n-butanol:acetic acid:water 3:1:1(by volume) mixture followed by ninhydrin stain. Calculate the R_f for each of the substances found on their TLC plate and then identify the substance using the reference table below.

Substance	Distance Traveled from Baseline (cm)	R_f	Identity of Substance
Solvent	6.0		
A	1.5	0.25	Glutamine
B	3.6	0.60	Tryptophan

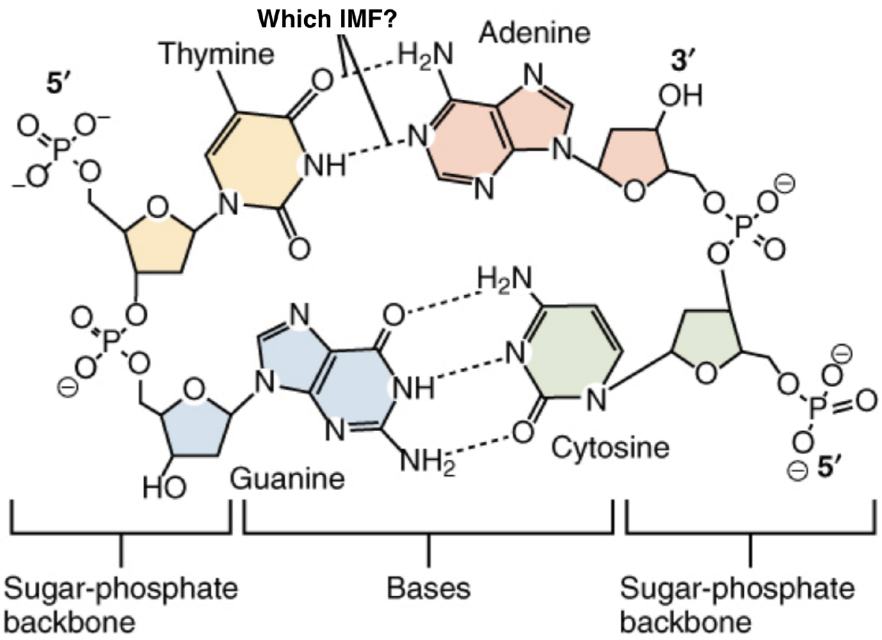
Retention Factors of Glutamine, Glutamic Acid, Tyrosine, Leucine, and Tryptophan in Silica Stationary Phase TLC with n-Butanol:Acetic Acid:Water 3:1:1(by volume) Mobile phase

Amino Acid	R_f Range
Glutamine	0.25-0.29
Glutamic Acid	0.31-0.35
Tyrosine	0.44-0.55
Tryptophan	0.57-0.61
Leucine	0.61-0.71

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

Hydrogen bonds hold together the two strands of DNA. By looking at the image, one can identify two of the three electronegative atoms capable of forming hydrogen bonds (nitrogen and oxygen) interacting with partially positive hydrogen atoms in the opposite DNA strand.

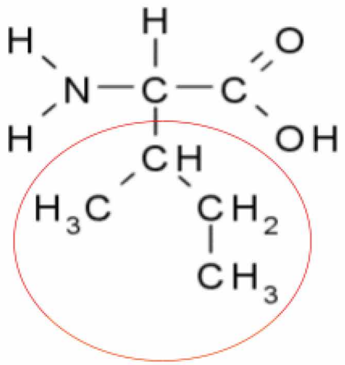
CLOSURE

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

DNA is a water soluble molecule. By looking at the structure in the image, one can see that there are ionic components of the DNA molecule as well as polar and hydrogen bonding-capable components. All of these things increase interactions with water molecules that lead to DNA's solubility in water.

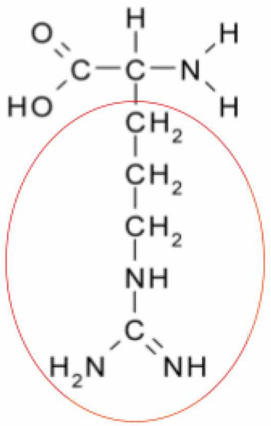
3. Protein folding is another application of intermolecular forces in biology. Dispersion forces, dipole-dipole interactions, hydrogen bonding, and even ion-dipole interactions all play key roles in determining how proteins fold into the extraordinary number of shapes that they do. Different amino acids 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



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Arginine is best suited for the area of a protein most exposed to water. As one can see, isoleucine's side chain only contains nonpolar bonds between carbon and hydrogen. Such a composition does not interact favorably with polar water molecules. Arginine, however, is not only polar due to the presence of nitrogen—carbon and nitrogen—hydrogen bonds, but is also capable of hydrogen bonding. These characteristics increase the interactions that the arginine side chain has with water molecules.