Intermolecular Forces 4.5

There are many physical properties that demonstrate the existence of **intermolecular forces**. For example, you would not want a raincoat made from untreated cotton. Cotton is very good at absorbing water and easily gets wet. The molecules that make up cotton can form many intermolecular attractions with water molecules. On the other hand, rubber or plastic materials do not absorb water because there is little intermolecular attraction between water molecules and the molecules of the rubber or plastic. A simple property like "wetting" depends to a large extent on intermolecular forces (**Figure 1**). Rubber or plastic may not be desirable materials to wear, but you could still use cotton if it is treated with a water repellent—a coating that has little attraction to water molecules. The development of water repellents requires a good knowledge of intermolecular forces.

Some bugs, like water spiders, walk on water and trees move water up large distances from the ground to the tops of the trees (**Figure 2**). Surface tension and capillary action are directly related to intermolecular attractions between molecules. In this section, we will look at these and other properties in terms of various intermolecular forces.

In 1873 Johannes van der Waals suggested that the deviations from the ideal gas law arose because the molecules of a gas have a small but definite volume and the molecules exert forces on each other. These forces are often simply referred to as van der Waals forces. It is now known that in many substances van der Waals forces are actually a combination of many types of intermolecular forces including, for example, *dipole–dipole forces* and *London forces*. Later, the concept of *hydrogen bonding* was created to explain anomalous (unexpected) properties of certain liquids and solids. In general, intermolecular forces are considerably weaker than the covalent bonds inside a molecule.

Intermolecular forces are much weaker than covalent bonds. As an approximate comparison, if covalent bonds are assigned a strength of about 100, then intermolecular forces are generally 0.001 to 15.

The evidence for this comparison comes primarily from experiments that measure bond energies. For example, it takes much less energy to boil water (breaking intermolecular bonds) than it does to decompose water (breaking covalent bonds).

$$H_2O_{(l)} \rightarrow H_2O_{(g)} \qquad 41\ \text{kJ/mol}$$

$$H_2O_{(l)} \rightarrow H_{2(g)} + \frac{1}{2}O_{2(g)} \qquad 242\ \text{kJ/mol}$$

Dipole–Dipole Force

In the last section, you learned how to test a stream of liquid to see whether the molecules of the liquid are polar. You also learned how to predict whether a molecule was polar or nonpolar. (Recall that polar molecules have dipoles—oppositely charged ends.) Attraction between dipoles is called the **dipole–dipole force** and is thought to be due to a simultaneous attraction between a dipole and surrounding dipoles (**Figure 3**).

intermolecular force the force of attraction and repulsion between molecules

Figure 1
Cotton (left) will absorb a lot more water than polyester, because the molecules in cotton are better able to attract and hold water molecules.

Figure 2
Water is transported in thin, hollow tubes in the trunk and branches of a tree. This is accomplished by several processes, including capillary action. It is also crucial that water molecules attract each other to maintain a continuous column of water.

DID YOU KNOW?

Other van der Waals Forces

Scientists today recognize that there are other varieties of van der Waals forces, such as dipole-induced dipole (part of a group of forces called induction forces). Dipole–dipole forces are now considered to be a special case of multipolar forces that also include quadrupole (4 pole) effects.

dipole–dipole force a force of attraction between polar molecules

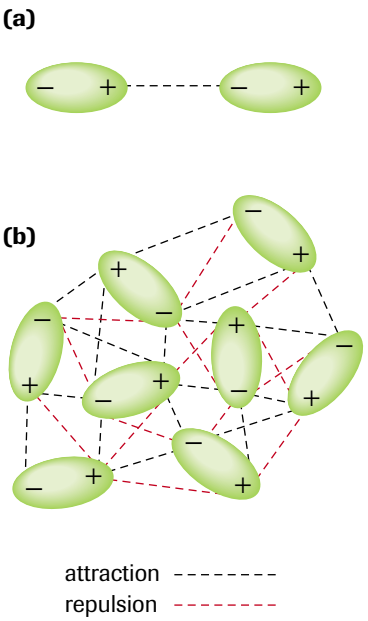


Figure 3
(a) Oppositely charged ends of polar molecules attract.
(b) In a liquid, polar molecules can move and rotate to maximize attractions and minimize repulsions. The net effect is a simultaneous attraction of dipoles.

- The dipole–dipole force is due to the simultaneous attraction of one dipole by its surrounding dipoles.
- The strength of the dipole–dipole force is dependent on the polarity of the molecule.

In the past you studied the effect of molecular polarity on solubility. The empirical generalization from that study was that "like dissolves like"; i.e., polar solutes dissolve in polar solvents and nonpolar solutes dissolve in nonpolar solvents. The most important polar solvent is, of course, water. Now we continue the study of intermolecular forces and their effect on several other physical properties of substances, such as boiling point, rate of evaporation, and surface tension. We will interpret these properties using intermolecular forces, but there are other factors that also affect these properties. To minimize this problem, we will try to compare simple, similar substances and do only qualitative comparisons. Before we try to tackle this problem we need to continue our look at kinds of intermolecular forces.

London Force

After repeated failures to find any pattern in physical properties like the boiling points of polar substances, Fritz London suggested that the van der Waals force was actually two forces—the dipole–dipole force and what we now call the **London force**. It was natural that London would describe the force of attraction between molecules (the intermolecular force) in the same way that he described the force of attraction within molecules (the intramolecular force). In both cases, London considered the electrostatic forces between protons and electrons. The difference is that for intramolecular forces the protons and electrons are in the same molecule, while for intermolecular forces the protons and electrons are in different molecules. London's analysis was actually more complicated than this, but for our purposes we only need one basic idea—an atomic nucleus not only attracts the electrons in its own molecule but also those in neighbouring molecules (**Figure 4**). London force is also called dispersion force or London dispersion force. This intermolecular force exists between all molecules. Logically, the strength of the London force depends on the number of electrons (and protons) in a molecule. The greater the number of electrons to be attracted to neighbouring nuclei, the stronger the resulting London force should be.

London force the simultaneous attraction of an electron by nuclei within a molecule and by nuclei in adjacent molecules

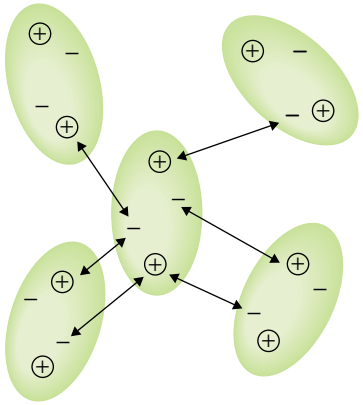


Figure 4
London force is an intermolecular attraction between all molecules. In this figure, only the attractions are shown.

- The London force is due to the simultaneous attraction of the electrons of one molecule by the positive nuclei in the surrounding molecules.
- The strength of the London force is directly related to the number of electrons in the molecule.

Using Dipole–Dipole and London Forces to Predict Boiling Points

Let's take a look at the boiling points of group 14 hydrogen compounds (**Table 1**). We would expect these molecules to be nonpolar, based on their four equivalent bonds and their tetrahedral shape.

In **Table 1**, you can see that as the number of electrons in the molecule increases (from 10 to 54), the boiling point increases (from −164°C to −52°C). The evidence presented in Table 1 supports the London theory, and provides a generalization for explaining and predicting the relative strength of London forces among molecules.

Table 1 The Boiling Points of Group 14 Hydrogen Compounds

Compound (at SATP)	Electrons	Boiling point (°C)
$CH_{4(g)}$	10	−164
$SiH_{4(g)}$	18	−112
$GeH_{4(g)}$	36	−89
$SnH_{4(g)}$	54	−52

Predicting Boiling Points

SAMPLE problem

1. ***Use London force theory to predict which of these alkanes has the highest boiling point—methane (CH_4), ethane (C_2H_6), propane (C_3H_8), or butane (C_4H_{10}).***

According to intermolecular-force theory, butane should have the highest boiling point. The reasoning behind this prediction is that all of these molecules are nonpolar, but butane has the most attractive London force, because it has the greatest number of electrons in its molecules.

This prediction is borne out by the evidence (**Table 2**).

Molecules with the same number of electrons (called **isoelectronic**) are predicted to have the same or nearly the same strengths for the London force of intermolecular attraction. Isoelectronic molecules help us to study intermolecular forces. For example, if one of two isoelectronic substances is polar and the other is nonpolar, then the polar molecule should have a higher boiling point, as shown in the following problem.

Table 2 Boiling Points of Alkanes

Alkane	Boiling point (°C)
methane	−162
ethane	−89
propane	−42
butane	−0.5

isoelectronic having the same number of electrons per atom, ion, or molecule

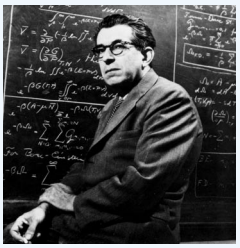
2. ***Consider the two isoelectronic substances, bromine (Br_2) and iodine monochloride (ICl). Based upon your knowledge of intermolecular forces, explain the difference in their boiling points (bromine, 59°C; iodine monochloride 97°C).***

Both bromine and iodine monochloride have 70 electrons per molecule (**Table 3**). Therefore, the strength of the London forces between molecules of each should be the same. Bromine is nonpolar and therefore has only London forces between its molecules. Iodine monochloride is polar, which means it has an extra dipole−dipole force between its molecules, in addition to London forces. This extra attraction among ICl molecules produces a higher boiling point.

Table 3 Isoelectronic Substances

Substance	Electrons	Boiling point (°C)
$Br_{2(l)}$	70	59
$ICl_{(l)}$	70	97

DID YOU KNOW?



This is the same London who worked with Sommerfeld and who Pauling indicated had made "the greatest single contribution to the chemist's conception of valence" since Lewis created the concept of the shared pair of electrons. The use of quantum mechanics to describe the covalent bond in the hydrogen molecule is the contribution referred to by Pauling.

SUMMARY — Predicting with Dipole–Dipole and London Forces

- Isoelectronic molecules have approximately the same strength of the London force.
- If all other factors are equal, then
  - the more polar the molecule, the stronger the dipole−dipole force and therefore, the higher the boiling point.
  - the greater the number of electrons per molecule, the stronger the London force and therefore, the higher the boiling point.
- You can explain and predict the relative boiling points of two substances if:
  - the London force is the same, but the dipole-dipole force is different.

— the dipole–dipole force is the same, but the London force is different.

— the influence of both the London force and the dipole–dipole force are in the same direction; e.g., both are tending to increase the boiling point of one of the chemicals.

- You cannot explain and predict with any certainty the relative boiling points of two chemicals if:

  — one of the substances has a stronger dipole–dipole force and the other substance has a stronger London force.

Practice

Understanding Concepts

1. Using London forces and dipole–dipole forces, state the kind of intermolecular force(s) present between molecules of the following substances:
   (a) water (solvent)
   (b) carbon dioxide (dry ice)
   (c) ethane (in natural gas)
   (d) ethanol (beverage alcohol)
   (e) ammonia (cleaning agent)
   (f) iodine (disinfectant)
2. Which of the following pure substances has stronger dipole–dipole forces than the other? Provide your reasoning.
   (a) hydrogen chloride or hydrogen fluoride
   (b) chloromethane or iodomethane
   (c) nitrogen tribromide or ammonia
   (d) water or hydrogen sulfide
3. Based upon London force theory, which of the following pure substances has the stronger London forces? Provide your reasoning.
   (a) methane or ethane
   (b) oxygen or nitrogen
   (c) sulfur dioxide or nitrogen dioxide
   (d) methane or ammonia
4. Based upon dipole–dipole and London forces, predict which substance in the following pairs has the higher boiling point. Provide your reasoning.
   (a) beryllium difluoride or oxygen difluoride
   (b) chloromethane or ethane
5. Why is it difficult to predict whether NF_3 or Cl_2O has the higher boiling point?

Applying Inquiry Skills

6. A common method in science is to gather or obtain experimental information and look for patterns. This is common when an area of study is relatively new and few generalizations exist. Analyze the information in **Table 4** to produce some possible patterns and then interpret as many patterns as possible using your knowledge of molecules and intermolecular forces.
7. Write an experimental design to test the ability of the theory and rules for the dipole–dipole force or the London force to predict the trend in melting points of several related substances. What are some possible complications with this proposed experiment?

Extension

8. Using a chemical reference, look up the boiling points for the substances in questions 4 and 5. Evaluate your predictions.

LAB EXERCISE 4.5.1

Boiling Points and Intermolecular Forces (p. 278)
This lab exercise shows some successes and some failures of the London force and dipole–dipole theories to predict and explain boiling points.

Table 4 Boiling Points of Hydrocarbons

Compound	Formula	Boiling point (°C)
ethane	C_2H_6	−89
ethene	C_2H_4	−104
ethyne	C_2H_2	−84
propane	C_3H_8	−42
propene	C_3H_6	−47
propyne	C_3H_4	−23
butane	C_4H_{10}	−0.5
1-butene	C_4H_8	−6
1-butyne	C_4H_6	8
pentane	C_5H_{12}	36
1-pentene	C_5H_{10}	30
1-pentyne	C_5H_8	40
hexane	C_6H_{14}	69
1-hexene	C_6H_{12}	63
1-hexyne	C_6H_{10}	71

Hydrogen Bonding

The unexpectedly high boiling points of hydrogen compounds of nitrogen (ammonia), oxygen (water), and fluorine (hydrogen fluoride) compared to those of hydrogen compounds of other elements in the same groups is evidence that some other effect in addition to dipole–dipole and London forces exists. Chemists have found that this behaviour is generalized to compounds where a hydrogen atom is bonded to a highly electronegative atom with a lone pair of electrons, such as nitrogen, oxygen, or fluorine (**Figure 5**). The explanation is that of **hydrogen bonding**. This process is not unlike a covalent bond, but in this case a proton is being shared between two pairs of electrons, rather than a pair of electrons being shared between protons.

The hydrogen bond was an extension of Lewis theory in 1920. The special properties of water and some other hydrogen-containing compounds required a new explanation. Maurice Huggins, a graduate student of Lewis, and two colleagues (Wendell Latimer and Worth Rodebush) devised the idea of a bond where hydrogen could be shared between some atoms like nitrogen, oxygen, and fluorine in two different molecules. This requires one of the two atoms to have a lone pair of electrons. Lewis referred to this as "a most important addition to my theory."

Additional evidence for hydrogen bonding can be obtained by looking at energy changes associated with the formation of hydrogen bonds. Recall that endothermic and exothermic reactions are explained by the difference between the energy absorbed to break bonds in the reactants and the energy released when new bonds in the products are formed. For example, in the exothermic formation of water from its elements, more energy is released in forming the new O−H bonds than is required to break H−H and O=O bonds. In a sample of glycerol, you would expect some hydrogen bonding between glycerol molecules. However, these molecules are rather bulky and this limits the number of possible hydrogen bonds. If water is mixed with glycerol, additional

hydrogen bonding the attraction of hydrogen atoms bonded to N, O, or F atoms to a lone pair of electrons of N, O, or F atoms in adjacent molecules

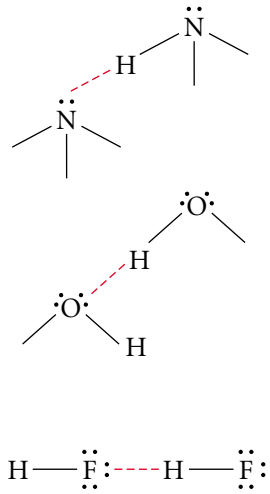
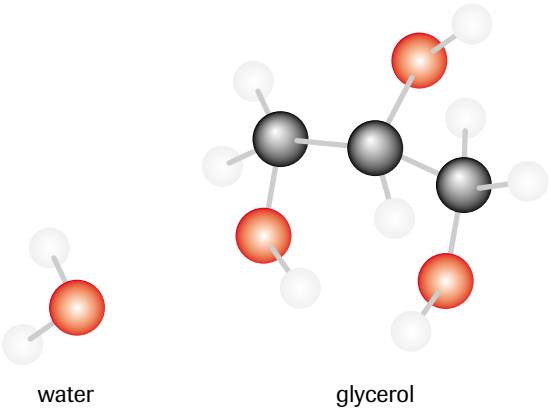


Figure 5
A hydrogen bond (--) occurs when a hydrogen atom bonded to a strongly electronegative atom is attracted to a lone pair of electrons in an adjacent molecule.



hydrogen bonds should be possible. The small size of the water molecule should make it possible for water to form many hydrogen bonds with the glycerol molecules. Experimentally, you find that mixing water with glycerol is an exothermic process.

The theory of hydrogen bonding is necessary to explain the functions of biologically important molecules. Recall that proteins are polymers of amino acids, and that amino acids have $-NH_2$ and $-COOH$ functional groups, both of which fulfill the conditions for hydrogen bonding. Similarly, the double helix of the DNA molecule owes its unique structure largely to hydrogen bonding. The central bonds that hold the double helix together are not covalent (**Figure 6**). If the helix were held together by covalent bonds, the DNA molecule would not be able to unravel and replicate.

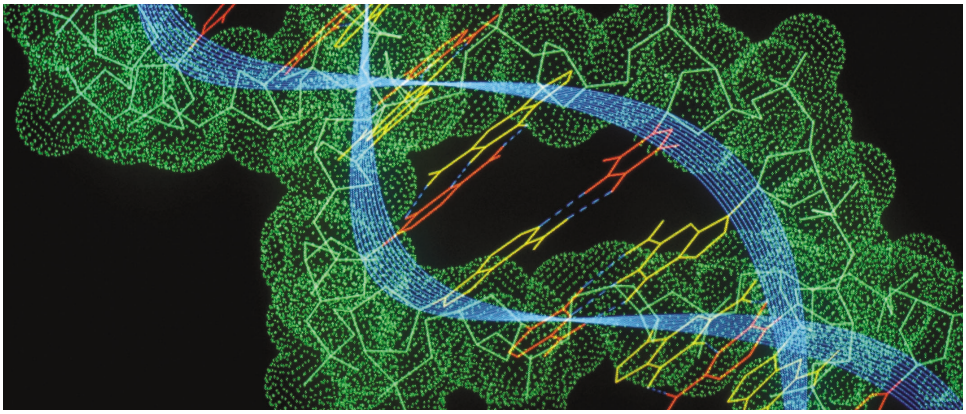


Figure 6
According to Francis Crick, co-discoverer with James D. Watson of the DNA structure, "If you want to understand function, study structure." Hydrogen bonding (blue dashes) explains the shape and function of the DNA molecule. The interior of the double helix is cross-linked by hydrogen bonds.

INVESTIGATION 4.5.1

Hydrogen Bonding (p. 278)
Mixing liquids provides evidence for hydrogen bonding.

Other Physical Properties of Liquids

Liquids have a variety of physical properties that can be explained by intermolecular forces. As you have seen, boiling point is a property often used in the discussion of intermolecular forces. Comparing boiling points provides a relatively simple comparison of intermolecular forces in liquids, if we assume that the gases produced have essentially no intermolecular forces between their molecules. What about some other properties of liquids, such as surface tension, shape of a meniscus, volatility, and ability to "wet" other substances? Surface tension is pretty important for water insects (**Figure 7**). The surface tension on a liquid is like an elastic skin. Molecules within a liquid are attracted by molecules on all sides, but molecules right at the surface are only attracted downward and sideways (**Figure 8**). This means that the liquid tends to stay together. Not surprisingly, substances containing molecules with stronger intermolecular forces have higher surface tensions. Water is a good example—it has one of the highest surface tensions of all liquids.

Figure 7
The weight of the water boatman is not enough to overcome the intermolecular forces between the water molecules. This would be like you walking on a trampoline. The fabric of the trampoline is strong enough to support your weight.

The shape of the meniscus of a liquid and capillary action in a narrow tube are both thought to be due to intermolecular forces. In both cases, two intermolecular attractions need to be considered—the attraction between like molecules (called cohesion) and the attraction between unlike molecules (called adhesion). Both cohesion and adhesion are intermolecular attractions. If you compare two very different liquids such as water and mercury (**Figure 9**), water rises in a narrow tube, but mercury does not. The adhesion between the water and the glass is thought to be greater than the cohesion between the water molecules. In a sense, the water is pulled up the tube by the intermolecular forces between the water molecules and the glass. Notice that this also produces a concave (curved downward) meniscus. For mercury, it is the opposite. The cohesion between the mercury atoms is greater than the adhesion of mercury to glass. Mercury atoms tend to stay together, which also explains the convex (curved upward) meniscus.

TRY THIS activity — Floating Pins

Materials: beaker or glass; water; several other different liquids; dishwashing detergent; straight pin; tweezers; toothpick

- Make sure the straight pin is clean and dry.
- Using clean tweezers, carefully place the pin in a horizontal position on the surface of each liquid, one at a time. Wash and dry the pin between tests.
  (a) What happens for each liquid? Why?
- Using tweezers, carefully place the pin vertically into the surface of the water. Try both ends of the pin.
  (b) What happens this time? Why do you think the result is different than before?
- Place the pin horizontally onto the surface of water. Using a toothpick, add a small quantity of dish detergent to the water surface away from the pin.
  (c) Describe and explain what happens.

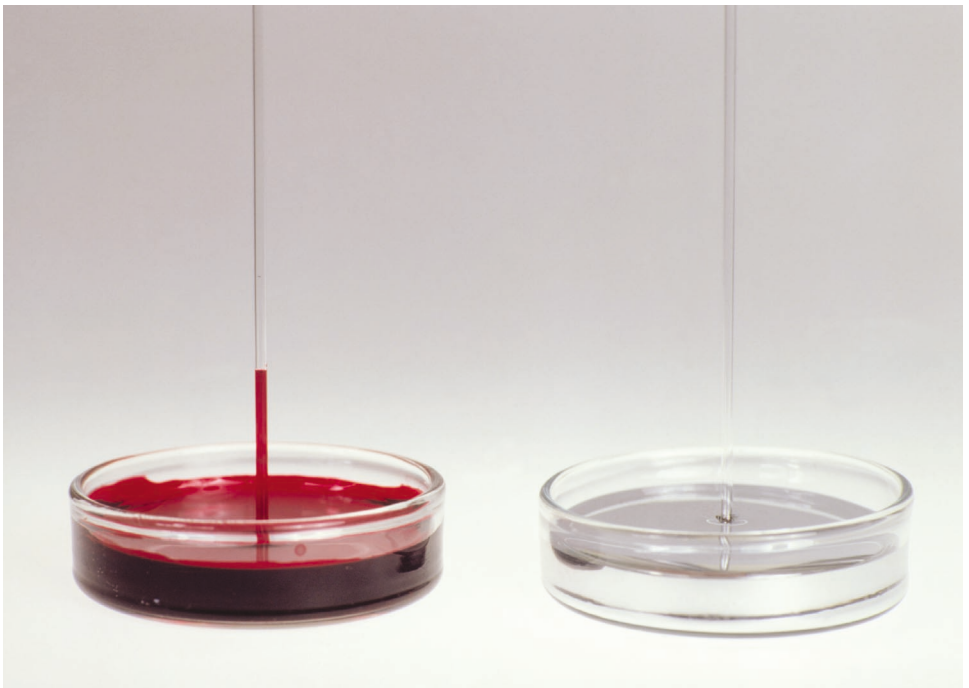


Figure 9
Capillary action is the movement of a liquid up a narrow tube. For water, capillary action is very noticeable; for mercury, it is nonexistent. (The water is coloured to make it more visible.)

We tend to think of water as a "normal" substance because it is so common and familiar to us. However, compared with other substances, water has some unusual properties. For example, the lower density of the solid form (ice) compared to the liquid form (water) is uncommon for a pure substance. Experimentally, the structure of ice is an open hexagonal network (**Figure 10**). Hydrogen bonding and the shape of the water molecule are believed to be responsible for this arrangement. Scientists have discovered that atoms and molecules can be trapped inside this hexagonal cage of water molecules. One of the more interesting and potentially valuable discoveries is the presence of large quantities of ice containing methane molecules on the ocean floor in the Arctic and around the globe (**Figure 11**). These deposits could be a vast new resource of natural gas in the future.

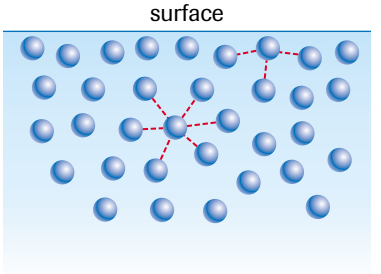


Figure 8
The intermolecular forces on a molecule inside a liquid are relatively balanced. The forces on a molecule right at the surface are not balanced—the net pull is toward the centre.

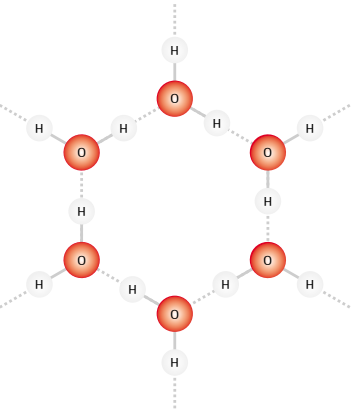


Figure 10
In ice and snowflakes, hydrogen bonds between water molecules result in an open hexagonal structure instead of a more compact structure with a higher density. The dashed lines represents the hydrogen bonds.

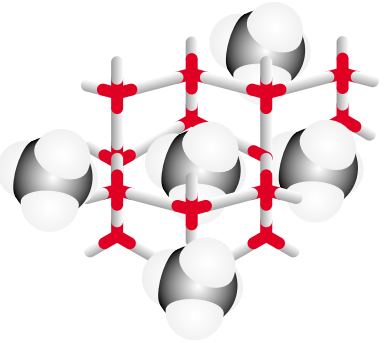


Figure 11
The very high pressures and low temperatures produce ice with methane trapped inside. The source of the methane is believed to be the decay of organic sediment deposited over a long time.

DID YOU KNOW?

Magic Sand

Ordinary sand is composed of silicates. Water can form intermolecular bonds with the oxygen atoms in the silicates and therefore "wet" the sand. Magic sand is sand coated with a nonpolar substance that cannot form intermolecular bonds with water molecules. Magic sand cannot be wetted. It forms shapes underwater and is completely dry when it is removed.

Figure 12

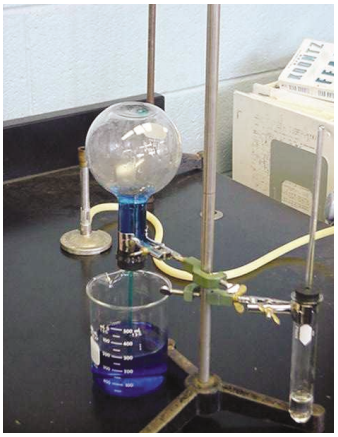


Figure 13

Practice

Understanding Concepts

9. For each of the following molecular compounds, hydrogen bonds contribute to the attraction between molecules. Draw a Lewis diagram using a dashed line to represent a hydrogen bond between two molecules of the substance.
(a) hydrogen peroxide, $H_2O_{2(l)}$ (disinfectant)
(b) hydrogen fluoride, $HF_{(l)}$ (aqueous solution etches glass)
(c) ethanol, $C_2H_5OH_{(l)}$ (beverage alcohol)
(d) ammonia, $NH_{3(l)}$ (anhydrous ammonia for fertilizer)

10. (a) Refer to or construct a graph of the evidence from Lab Exercise 4.5.1. Extrapolate the group 15 and 16 lines to estimate the boiling points of water and ammonia if they followed the trend of the rest of their family members.
(b) Approximately how many degrees higher are the actual boiling points for water and ammonia compared to your estimate in (a)?
(c) Explain why the actual boiling points are significantly higher for both water and ammonia.
(d) Propose an explanation why the difference from (b) is much greater for water than for ammonia.

11. Water beads on the surface of a freshly waxed car hood. Use your knowledge of intermolecular forces to explain this observation.

12. A lava lamp is a mixture of two liquids with a light bulb at the bottom to provide heat and light (**Figure 12**). What interpretations can you make about the liquids, intermolecular forces, and the operation of the lamp?

Applying Inquiry Skills

13. To gather evidence for the existence of hydrogen bonding in a series of chemicals, what variables must be controlled?

14. (a) Design an experiment to determine the volatility (rate of evaporation) of several liquids. Be sure to include variables.
(b) Suggest some liquids to be used in this experiment. Predict the results. Explain your reasoning.

15. **Question**
What is the solubility (high or low) of ammonia in water?

Prediction/Hypothesis
(a) Write a prediction, including your reasoning.

Experimental Design
Some water is squirted into ammonia gas in a Florence flask. Another tube is available for drawing water up into the flask.

Evidence
The water starts to move slowly up the tube and then suddenly flows into the upper flask like a fountain (**Figure 13**).

Analysis
(b) Answer the question, based upon the evidence gathered.

Evaluation
(c) Assuming that you have confidence in the evidence presented, evaluate the prediction and the reasoning used to make the prediction.

Making Connections

16. Wetting agents are very important in agriculture and other industries. What are wetting agents? Where are they used and for what purposes? Briefly explain how the function of wetting agents relates to the principles of intermolecular forces.

GO www.science.nelson.com

17. In 1966 Soviet scientists claimed to have discovered a new form of water, called polywater. The story of polywater is an interesting example of how people, including scientists, want to believe in a new, exciting discovery even if the evidence is incomplete. Write a brief report about polywater, including how it is supposedly formed, some of its claimed properties, the explanation in terms of intermolecular forces, and the final evaluation of the evidence (specifically, the flaws).

GO www.science.nelson.com

SUMMARY *Intermolecular Forces*

- Intermolecular forces, like all bonds, are electrostatic—they involve the attraction of positive and negative charges.
- In this section we considered three intermolecular forces—London, dipole–dipole, and hydrogen bonding.
- All molecules attract each other through the London force—the simultaneous attraction of electrons and nuclei in adjacent molecules.
- Dipole–dipole force exists between polar molecules—the simultaneous attraction of a dipole of one molecule for adjacent dipoles.
- Hydrogen bonding exists when hydrogen atoms are bonded to highly electronegative atoms like N, O, and F—the hydrogen is simultaneously attracted to a pair of electrons on the N, O, or F atom of an adjacent molecule.
- Intermolecular forces affect the melting point, boiling point, capillary action, surface tension, volatility, and solubility of substances.

Current Research—Intermolecular Forces

In almost any area of science today, the experimental work runs parallel to the theoretical work and there is constant interplay between the two areas. In Canada there are several theorists whose research teams examine the forces between atoms and molecules to increase our understanding of physical and chemical properties. One such individual is Dr. Robert LeRoy (**Figure 14**), currently working in theoretical chemical physics at the University of Waterloo.

Dr. LeRoy's interest is intermolecular forces. He uses quantum mechanics and computer models to define and analyze the basic forces between atoms and molecules. Early in his career, Dr. LeRoy developed a technique for mathematically defining a radius of a small molecule, now known as the LeRoy radius. This established a boundary. Within the boundary, intramolecular bonding is important, and beyond the boundary, intermolecular forces predominate. In his work, the study of atomic and molecular spectra (called spectroscopy) plays a crucial role. Measurements from spectroscopy help theoreticians develop better models and theories for explaining molecular structure. Computer programs that Dr. LeRoy has developed for the purpose of converting experimental evidence to information on forces, shape, and structure are free, and are now routinely used around the world.

It is important not to assume that forces and structures are well established. Our knowledge of bonding and structure becomes more and more scanty and unreliable for larger structures. A huge amount of research remains to be done if we are ever to be able to describe bonding and structure very accurately for even microscopic amounts of

Figure 14
Dr. Robert J. LeRoy and his research team study intermolecular forces and the behaviour of small molecules and molecular clusters; develop methods to simulate and analyze the decomposition of small molecules; and create computer models to simulate and predict molecular properties. Visit Dr. LeRoy at http://leroy.uwaterloo.ca.

complex substances. Dr. LeRoy states "... except for the simplest systems, our knowledge of (interactions between molecules) is fairly primitive...." A classic example is our understanding of the structure and activity of proteins—the stuff of life. We know the composition of many proteins quite precisely and the structure can be experimentally determined, but the structure of these large molecules depends on how bonding folds and shapes the chains and branches. How a protein behaves and what it does depends specifically on its precise shape and structure, and that is something scientists often state is "not well understood."

Section 4.5 Questions

Understanding Concepts

1. All molecular compounds may have London, dipole–dipole, and hydrogen-bonding intermolecular forces affecting their physical and chemical properties. Indicate which intermolecular forces contribute to the attraction between molecules in each of the following classes of organic compounds:
   (a) hydrocarbon; e.g., pentane, $C_5H_{12(l)}$ (in gasoline)
   (b) alcohol; e.g., 2-propanol, $CH_3CHOHCH_{3(l)}$ (rubbing alcohol)
   (c) ether; e.g., dimethylether, $CH_3OCH_{3(g)}$ (polymerization catalyst)
   (d) carboxylic acid; e.g., acetic acid, $CH_3COOH_{(l)}$ (in vinegar)
   (e) ester; e.g., ethylbenzoate, $C_6H_5COOC_2H_{5(l)}$ (cherry flavour)
   (f) amine; e.g., dimethylamine, $CH_3NHCH_{3(g)}$ (depilatory agent)
   (g) amide; e.g., ethanamide, $CH_3CONH_{2(s)}$ (lacquers)
   (h) aldehyde; e.g., methanal, $HCHO_{(g)}$ (corrosion inhibitor)
   (i) ketone; e.g., acetone, $(CH_3)_2CO_{(l)}$ (varnish solvent)
2. Use Lewis structures and hydrogen bonds to explain the very high solubility of ammonia in water.
3. Predict the solubility of the following organic compounds in water as low (negligible), medium, or high. Provide your reasoning.
   (a) 2-chloropropane, $C_3H_7Cl_{(l)}$ (solvent)
   (b) 1-propanol, $C_3H_7OH_{(l)}$ (brake fluids)
   (c) propanone, $(CH_3)_2CO_{(l)}$ (cleaning precision equipment)
   (d) propane, $C_3H_{8(g)}$ (gas barbecue fuel)
4. For each of the following pairs of chemicals, which one is predicted to have the stronger intermolecular attraction? Provide your reasoning.
   (a) chlorine or bromine
   (b) fluorine or hydrogen chloride
   (c) methane or ammonia
   (d) water or hydrogen sulfide
   (e) silicon tetrahydride or methane
   (f) chloromethane or ethanol
5. Which liquid, propane (C_3H_8) or ethanol (C_2H_5OH), would have the greater surface tension? Justify your answer.
6. In cold climates, outside water pipes, such as underground sprinkler systems, need to have the water removed before it freezes. What might happen if water freezes in the pipes? Explain your answer.
7. A glass can be filled slightly above the brim with water without the water running down the outside. Explain why the water does not overflow even though some of it is above the glass rim.
8. Explain briefly what the "LeRoy radius" of a molecule represents.

Applying Inquiry Skills

9. Design an experiment to determine whether or not hydrogen bonding has an effect on the surface tension of a liquid. Clearly indicate the variables in this experiment.
10. Critique the following experimental design.
    The relative strength of intermolecular forces in a variety of liquids is determined by measuring the height to which the liquids rise in a variety of capillary tubes.

Making Connections

11. Some vitamins are water soluble (e.g., B series and C), while some are fat soluble (e.g., A, D, E, and K).
    (a) What can you infer about the polarity of these chemicals?
    (b) Find and draw the structure of at least one of the water-soluble and one of the fat-soluble vitamins.
    (c) When taking vitamins naturally or as supplements, what dietary requirements are necessary to make sure that the vitamins are used by the body?
    (d) More of a vitamin is not necessarily better. Why can you take a large quantity of vitamin C with no harm (other than the cost), but an excess of vitamin E can be dangerous?

    GO www.science.nelson.com

12. Many of the new materials that are being invented for specific purposes show an understanding of structure and bonding. One candidate that has been suggested as a future product is commonly known as the "fuzzyball," $C_{60}F_{60(s)}$. What is the structure of this molecule? What use is proposed for this substance? Explain this use in terms of intermolecular forces.

    GO www.science.nelson.com

13. People who wear contact lenses know that there are hard and soft contact lenses. The polymers used in each type of lens are specifically chosen for their properties. What is the property that largely determines whether the lens is a hard or soft lens? Write a brief explanation using your knowledge of intermolecular forces.

GO www.science.nelson.com

14. Plastic cling wrap is widely used in our society. Why does it cling well to smooth glass and ceramics, but not to metals? Describe the controversial social issue associated with the use of this plastic wrap. How are intermolecular forces involved in starting the process that leads to this controversy?

GO www.science.nelson.com

Extensions

15. The London force is affected by more than just the number of electrons. What other variable(s) affect(s) the strength of the London force?

16. (a) Draw a bar graph with the temperature in kelvin on the vertical axis and the three isoelectronic compounds listed below on the horizontal axis. For each compound draw a vertical bar from 0 K to its boiling point: propane (−42°C), fluoroethane (−38°C), and ethanol (78°C).

(b) Divide each of the three bar graphs into the approximate component for the intermolecular force involved. (Assume that the London force is the same for each chemical and that the dipole−dipole force is the same for the two polar molecules.)

(c) Based upon the proportional components for the three possible intermolecular forces, order the relative strength of these forces.