Skills to Develop

• Define *phase*.
• Identify the types of interactions between molecules.

A phase is a certain form of matter that includes a specific set of physical properties. That is, the atoms, the molecules, or the ions that make up the phase do so in a consistent manner throughout the phase. Science recognizes three stable phases: the *solid phase*, in which individual particles can be thought of as in contact and held in place; the *liquid phase*, in which individual particles are in contact but moving with respect to each other; and the *gas phase*, in which individual particles are separated from each other by relatively large distances. Not all substances will readily exhibit all phases. For example, carbon dioxide does not exhibit a liquid phase unless the pressure is greater than about six times normal atmospheric pressure. Other substances, especially complex organic molecules, may decompose at higher temperatures, rather than becoming a liquid or a gas.

For many substances, there are different arrangements the particles can take in the solid phase, depending on temperature and pressure.

Which phase a substance adopts depends on the pressure and the temperature it experiences. Of these two conditions, temperature variations are more obviously related to the phase of a substance. When it is very cold, H\(_2\)O exists in the solid form as ice. When it is warmer, the liquid phase of H\(_2\)O is present. At even higher temperatures, H\(_2\)O boils and becomes steam.

Pressure changes can also affect the presence of a particular phase (as we indicated for carbon dioxide), but its effects are less obvious most of the time. We will mostly focus on the temperature effects on phases, mentioning pressure effects only when they are important. Most chemical substances follow the same pattern of phases when going from a low temperature to a high temperature: the solid phase, then the liquid phase, and then the gas phase. However, the temperatures at which these phases are present differ for all substances and can be rather extreme. Table 1 shows the temperature ranges for solid, liquid, and gas phases for three substances. As you can see, there is extreme variability in the temperature ranges.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Solid Phase Below</th>
<th>Liquid Phase Above</th>
<th>Gas Phase Above</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen (H(_2))</td>
<td>−259°C</td>
<td>−259°C</td>
<td>−253°C</td>
</tr>
<tr>
<td>water (H(_2)O)</td>
<td>0°C</td>
<td>0°C</td>
<td>100°C</td>
</tr>
<tr>
<td>sodium chloride (NaCl)</td>
<td>801°C</td>
<td>801°C</td>
<td>1413°C</td>
</tr>
</tbody>
</table>

The *melting point* of a substance is the temperature that separates a solid and a liquid. The *boiling point* of a substance is the temperature that separates a liquid and a gas.

What accounts for this variability? Why do some substances become liquids at very low temperatures, while others require very high temperatures before they become liquids? It all depends on the strength of the intermolecular forces.
interactions between the particles of substances. (Although ionic compounds are not composed of discrete molecules, we will still use the term *intermolecular* to include interactions between the ions in such compounds.) Substances that experience strong intermolecular interactions require higher temperatures to become liquids and, finally, gases. Substances that experience weak intermolecular interactions do not need much energy (as measured by temperature) to become liquids and gases and will exhibit these phases at lower temperatures.

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**Covalent Network Substances**

Substances with the highest melting and boiling points have covalent network bonding. This substances are held together by covalent bonds. In these substances, all the atoms in a sample are covalently bonded to other atoms. In effect, the entire sample is essentially one large molecule so the term intermolecular (or between molecules) is often not used for these interactions (since there are not multiple, discreet molecules). Many of these substances are solid over a large temperature range because it takes a lot of energy to disrupt all the covalent bonds at once. One example of a substance that shows covalent network bonding is diamond (Figure 1), which is a form of pure carbon. At temperatures over 3,500°C, diamond finally vaporizes into gas-phase atoms.

![Figure 1: Diamond. Diamond, a form of pure carbon, has covalent network bonding. It takes a very high temperature—over 3,500°C—for diamond to leave the solid state. Source: Photo © Thinkstock](image)

**Ionic Compounds**

The strongest force between any two particles is the *ionic bond*, in which two ions of opposing charge are attracted to each other. While not ions are not technically molecules, ionic interactions between particles are usually classified as a type of intermolecular interaction. Substances that contain ionic interactions are relatively strongly held together, so these substances typically have high melting and boiling points. Sodium chloride (Figure 2) is an example of a substance whose particles experience ionic interactions (Table 1).
Covalent Compounds (also called Molecular Compounds)

Many substances that experience covalent bonding exist as discrete molecules. Very strong covalent bonds within the molecule hold the atoms together to make each molecule. It is not, however, covalent bonds that are responsible for bringing the separate molecules together to make liquids and solids. Instead, there are three types of intermolecular interactions that exist between molecules that hold liquids and solids together. All thee types result from the attractions between partial opposite charges and are weaker than covalent or ionic bonds.

In many molecules, the electrons that are shared in a covalent bond are not shared equally among the two atoms in the bond. Typically, one of the atoms attracts the electrons more strongly than the other, leading to an unequal sharing of electrons in the bond. This idea is illustrated in Figure \(\PageIndex{3}\), which shows a diagram of the covalent bond in hydrogen fluoride (HF). The fluorine atom attracts the electrons in the bond more than the hydrogen atom does. An arrow is often used to show in which direction the electrons are pulled. The result is an unequal distribution of electrons in the bond, favoring the fluorine side of the covalent bond. Because of this unequal distribution, the fluorine side of the covalent bond actually takes on a partial negative charge (indicated by the \(\delta^−\) in Figure \(\PageIndex{3}\)), while the hydrogen side of the bond, being electron deficient, takes on a partial positive charge (indicated by the \(\delta^+\) in Figure \(\PageIndex{3}\)). A covalent bond that has an unequal sharing of electrons is called a polar covalent bond. (A covalent bond that has an equal sharing of electrons, as in a covalent bond with the same atom on each side, is called a nonpolar covalent bond.) A molecule with a net unequal distribution of electrons in its covalent bonds is a polar molecule. HF is an example of a polar molecule.

![Diagram of polar covalent bond](image)

**Figure \(\PageIndex{3}\):** Polar Covalent Bonds. The electrons in the HF molecule are not equally shared by the two atoms in the bond. Because the fluorine atom has nine protons in its nucleus, it attracts the negatively charged electrons in the bond more than the hydrogen atom does with its one proton in its nucleus. Thus, electrons are more strongly attracted to the fluorine atom, leading to an imbalance in the electron distribution between the atoms. The fluorine side of the bond picks up a partial overall negative charge (represented by the \(\delta^−\) in the diagram), while the hydrogen side of the bond has an overall partial positive charge (represented by the \(\delta^+\) in the diagram). Such a bond is called a polar covalent bond.

The charge separation in a polar covalent bond is not as extreme as is found in ionic compounds, but there is a related result: oppositely charged ends of different molecules will attract each other. These types of intermolecular interactions are called a dipole-dipole forces. Many molecules with polar covalent bonds experience dipole-dipole forces. The covalent bonds in some molecules are oriented in space symmetrically so that the bonds in the molecules cancel each other out. The individual bonds are polar, but the overall molecule is not polar; rather, the molecule is nonpolar due to its
overall symmetry. Such molecules experience little or no dipole-dipole interactions. Carbon dioxide (CO₂) and carbon tetrachloride (CCl₄) are examples of such molecules (Figure \(\PageIndex{4}\)).

**Figure \(\PageIndex{4}\):** Nonpolar Molecules. Although the individual bonds in both CO₂ and CCl₄ are polar, their effects cancel out because of the symmetrical distribution of the bonds in each molecule. As a result, such molecules experience little or no dipole-dipole interaction.

The H–F, O–H, and N–H bonds are strongly polar; in molecules that have these bonds, particularly strong dipole-dipole interactions (as strong as 10% of a true covalent bond) can occur. Because of this strong interaction, the term hydrogen bonding is used to describe this dipole-dipole interaction. The physical properties of water, which has two O–H bonds, are strongly affected by the presence of hydrogen bonding forces of attraction between water molecules. Figure \(\PageIndex{5}\) shows how molecules experiencing hydrogen bonding can interact. Hydrogen-bonding forces between water molecules (illustrated below as green dashes) are different than the covalent O-H bonds within a water molecule (illustrated as solid orange lines below). Hydrogen-bonding forces between molecules are stronger than dipole-dipole forces between molecules, but not as strong as covalent bonds within molecules.

**Figure \(\PageIndex{5}\):** Hydrogen-Bonding forces between Water Molecules (dashed green lines). The presence of hydrogen bonding in molecules like water can have a large impact on the physical properties of a substance.
Finally, there are forces between all molecules that are caused by electrons being in different places in a molecule at any one time, which sets up a temporary separation of charge that disappears almost as soon as it appears. These are very weak intermolecular interactions and are called dispersion forces (or London forces). (An alternate name is London-dispersion forces.) Molecules that experience no other type of intermolecular interaction will at least experience dispersion forces. Substances that experience only dispersion forces are typically soft in the solid phase and have relatively low melting points. Because dispersion forces are caused by the instantaneous distribution of electrons in a molecule, larger molecules with a large number of electrons can experience substantial dispersion forces. Examples include waxes, which are long hydrocarbon chains that are solids at room temperature because the molecules have so many electrons. The resulting dispersion forces between these molecules make them assume the solid phase at normal temperatures.

The phase that a substance adopts depends on the type and magnitude of the intermolecular interactions the particles of a substance experience. If the intermolecular interactions are relatively strong, then a large amount of energy—in terms of temperature—is necessary for a substance to change phases. If the intermolecular interactions are weak, a low temperature is all that is necessary to move a substance out of the solid phase.

Example \(\PageIndex{1}\): Intermolecular Forces

What intermolecular interactions besides dispersion forces, if any, exist in each substance? Are any of these substances solids at room temperature?

a. potassium chloride (KCl)

b. ethanol (C\(_2\)H\(_5\)OH)

c. bromine (Br\(_2\))

SOLUTION

a. Potassium chloride is composed of fully charged ions, so the intermolecular interactions in potassium chloride are ionic bonds. Because ionic bonds are strong, it might be expected that potassium chloride is a solid at room temperature.

b. Ethanol has a hydrogen atom attached to an oxygen atom, so it would experience hydrogen-bonding forces. It is also polar so has dipole-dipole forces, but the H-bonding forces are more significant. If the hydrogen bonding is strong enough, ethanol might be a solid at room temperature, but it is difficult to know for certain. (Ethanol is actually a liquid at room temperature.)

c. Elemental bromine has two bromine atoms covalently bonded to each other. Because the atoms on either side of the covalent bond are the same, the electrons in the covalent bond are shared equally, and the bond is a nonpolar covalent bond. Thus, diatomic bromine does not have any intermolecular forces other than dispersion forces. It is unlikely to be a solid at room temperature unless the dispersion forces are strong enough. Bromine is a liquid at room temperature.

Concept Review Exercise

1. What types of interactions might cause a gas to stick together as liquid or solid in:

   A. a covalent network substance
B. an ionic compound

C. a molecular compound?

**Answer**

1. A. covalent bonds (sharing of electrons)

B. ionic bonds (attraction between full opposite charges of ions)

C. All molecules will have London-dispersion forces, polar molecules will have dipole-dipole forces, and molecules with H directly bonded to N, O, or F will have hydrogen-bonding forces. (All three are attractions of partial opposite charges between molecules.)

**Key Takeaways**

- A phase is a form of matter that has the same physical properties throughout.
- Atoms and molecules interact with each other through various forces: ionic and covalent bonds, dipole-dipole interactions, hydrogen bonding, and dispersion forces.

**Exercises**

1. List the three common phases in the order you are likely to find them—from lowest temperature to highest temperature.

2. List these intermolecular interactions from weakest to strongest: London forces, hydrogen-bonding forces, dipole-dipole forces and ionic interactions.

3. What type of intermolecular interaction is predominate in each substance?
   a. ammonia (NH₃)
   b. sodium sulfate (Na₂SO₄)
   c. decane (C₁₀H₂₂)

4. Explain how a molecule like carbon dioxide (CO₂) can have polar covalent bonds but be nonpolar overall.

5. What types of intermolecular forces would exist in a sample of sulfur dichloride (SCl₂)?

6. What are some of the physical properties of substances that experience covalent network bonding?

7. What are some of the physical properties of substances that experience only dispersion forces?

   8. Heat is required to boil water (endothermic). What types of interactions are being overcome when water is boiled?
Answers

1. solid, liquid, and gas
2. London forces < dipole-dipole forces < hydrogen-bonding forces < ionic interactions
3. a. hydrogen-bonding forces predominate (but NH₃ also has dipole-dipole and London-dispersion); b. ionic interactions; c. dispersion forces (C-C and C-H bonds are not polar)
4. In linear carbon dioxide, the two covalent bonds are oriented on opposite sides (symmetrically) such that the polarity of each bond cancels out.
5. Sulfur dichloride has polar bonds, and because of the two lone pairs on the central sulfur atom, would have a bent geometry. The molecule is not symmetrical so one side is more negative and the other more positive. Like all molecules, it would have London-dispersion forces and, because it is polar, also dipole-dipole forces.
6. very hard, high melting point
7. low melting point and low boiling point
8. London-dispersion, dipole-dipole, and hydrogen-bonding forces of attraction are being overcome to separate the liquid water molecules into gaseous water molecules. (Covalent bonds bind oxygen to the two hydrogens in each water, but these are not broken when water is boiled.)