Having considered the bonding situation with hydrogen and helium, the simplest two elements, we can now move on to consider other elements and the types of molecules that they form. In this discussion, we begin with molecules made up of a single type of atom. More complex molecules made of atoms of multiple elements will be considered in the next and subsequent chapters. As the number of protons in the nucleus of an element’s atoms increases, from 1 in hydrogen to 10 in neon, we find dramatic changes in physical properties that correlate with whether the elemental form is discrete or continuous. The discrete forms are either monoatomic—meaning that they exist as separate atoms (such as He and Ne) with no covalent bonds between them (although they do interact via van der Waals interactions)—or diatomic molecules (such as H₂, N₂, O₂, and F₂), meaning that they exist as molecules that have only two atoms. The elements that exist as small molecules have very low melting points (the temperatures at which they change from a solid to a liquid) and low boiling points (the temperatures at which they change from a liquid to a gas). But don’t confuse these phase transitions with the breaking of a diatomic molecule into separate atoms. Phase transitions, which we will discuss in greater detail later, involve disruption of interactions between molecules (intermolecular forces), such as London dispersion forces, rather than interactions within molecules, that is, covalent bonds.

In contrast to the elements that form discrete molecules, the atoms of the other elements we are considering (that is Li, Be, B, C) interact with one another in a continuous manner. Rather than forming discrete molecules, these elements can form ensembles of atoms in which the number of atoms can range from the small (a few billion) to the astronomical (very, very large). Whether the materials are at the nano- or the macroscopic levels, the atoms in these ensembles are held together by bonds that are very difficult to break, like the bond in H–H. That is, a lot of energy must be put into the system to separate the component atoms. However, unlike hydrogen, the atoms that form these structures must form bonds with more than one other atom.

### Table 3.1 The First 10 Elements in Their Naturally Occurring Elemental State

<table>
<thead>
<tr>
<th>Elemental Form</th>
<th>H₂</th>
<th>He</th>
<th>Li</th>
<th>Be</th>
<th>B</th>
<th>C</th>
<th>N₂</th>
<th>O₂</th>
<th>F₂</th>
<th>Ne</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting Point (K)</td>
<td>13.81</td>
<td>0.95</td>
<td>453.65</td>
<td>1560</td>
<td>2348</td>
<td>3823</td>
<td>63.15</td>
<td>54.36</td>
<td>53.53</td>
<td>24.56</td>
</tr>
<tr>
<td>Boiling Point (K)</td>
<td>20.28</td>
<td>4.22</td>
<td>1615</td>
<td>2744</td>
<td>4273</td>
<td>4098</td>
<td>77.36</td>
<td>90.20</td>
<td>85.03</td>
<td>27.07</td>
</tr>
<tr>
<td>Bp-Mp (*) (K)</td>
<td>6.47</td>
<td>3.27</td>
<td>1161</td>
<td>1184</td>
<td>1925</td>
<td>275</td>
<td>14.21</td>
<td>35.84</td>
<td>31.5</td>
<td>2.51</td>
</tr>
<tr>
<td>Name</td>
<td>hydrogen</td>
<td>helium</td>
<td>lithium</td>
<td>beryllium</td>
<td>boron</td>
<td>carbon</td>
<td>nitrogen</td>
<td>oxygen</td>
<td>fluorine</td>
<td>neon</td>
</tr>
</tbody>
</table>

* boiling point minus melting point.

A consequence of this difference in organization is a dramatic increase in both the melting and boiling points compared to atomic (He, 

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1
Ne) and molecular (H₂, N₂, O₂, and F₂) species (Table 3.1). The reason is that when a substance changes from solid to liquid (at the melting point) the component particles have to be able to move relative to one another. When the substance changes from a liquid to a gas (at the boiling point) the particles have to separate entirely. Consequently the magnitude of the melting and boiling points gives us a relative estimate of how strongly the particles are held together in the solid and liquid states. As we have already seen temperature is a measure of the average kinetic energy of the molecules in a system. For elements that exist as discrete atoms or molecules the only forces that are holding these particles together are London dispersion forces, which are relatively weak compared to covalent bonds. In contrast, the elements that exist as extensive networks of atoms joined by bonds require much more energy to break as the material goes from solid to liquid to gas.

Questions to Ponder

1. Are all bonds the same?
2. What factors might influence bond strength?
3. Why are the properties of atoms and molecules different?

Questions to Answer

1. Where are the electrons in H₂ when the temperature is greater than 5000 K?
2. What would happen if you could form a He–He system with 3 electrons (instead of 4)?
3. What would a molecular-level picture of H₂ (g) look like?
4. What would a molecular-level picture of H (g) look like?
5. Where does the energy to break a bond come from?
6. Where does the energy released upon bond formation go?
7. The melting point of molecular hydrogen (H₂) is ~14 K (-259 °C). Draw a molecular level picture of what molecular hydrogen looks like below this temperature (as a solid). Why are the molecules of hydrogen sticking together?
8. The boiling point of molecular hydrogen (H₂) is ~20 K (-253 °C). Draw a molecular level picture of what molecular hydrogen looks like above this temperature (as a gas).
9. Molecular hydrogen dissociates at high temperatures (> 6000 K). Draw a picture of what you imagine this might look like. Why do you think it takes such a high temperature to bring about this change?