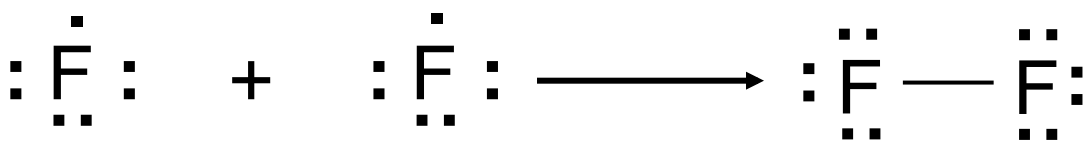
(Section 9.5) Covalent Bonding

- Covalent bonds involve two nonmetal atoms sharing one or more pairs of electrons.



- Covalent compounds usually exist as discreet molecules, but extended arrays of covalent bonds are possible.

The Octet Rule

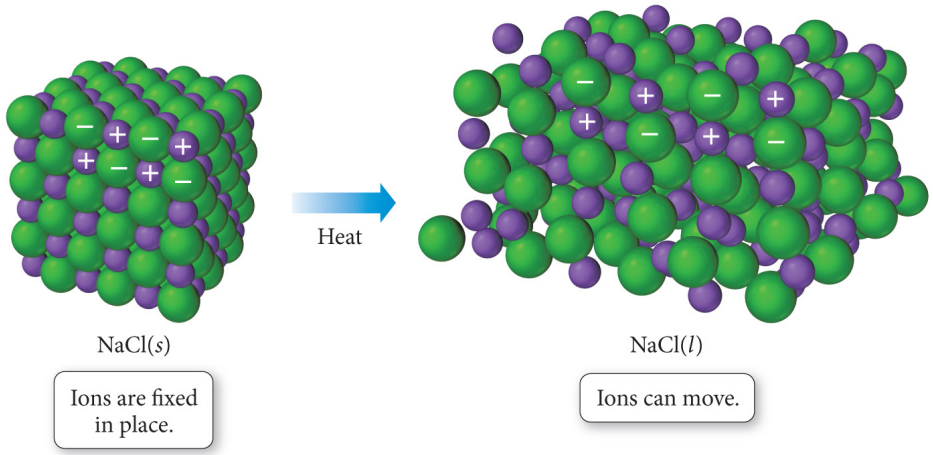
- Atoms will usually share enough electrons in order to be surrounded by **8 electrons**, since this gives a noble gas configuration.
  - except hydrogen: needs only 2 electrons to have same configuration as He.

Double and Triple Bonds

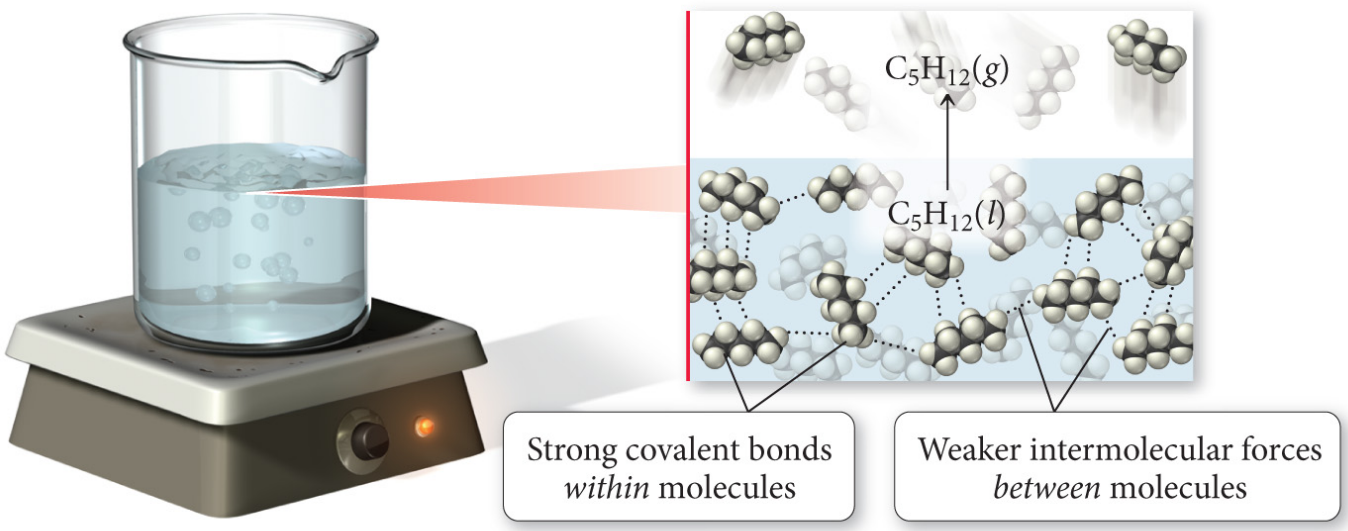
- Sometimes atoms need to share two pairs of electrons in order for each to have an octet.
- This is called a **double bond** – contains 4 electrons.
- A double bond is stronger (and shorter) than a single bond.

- In a **triple bond**, atoms share 3 pairs of electrons.
- A triple bond is even shorter and stronger than a double bond.

Comparing Covalent and Ionic Bonding



Molecular Compound



Comparing Covalent and Ionic Bonding

- The covalent bonds between atoms **within** molecules are strong, but a molecule is generally more weakly attracted to other molecules (Figure 9.5).
- So it is relatively easy to separate molecules and boiling points of molecular substances tend to be quite low.
- Melting or boiling ionic compounds involves overcoming the lattice energy, so they are are usually solids with very high m.p / b.p (*e.g.* NaCl melts at 800 °C).

- **Summary**: Ionic compounds tend to melt and boil at much higher temperatures than molecular compounds.

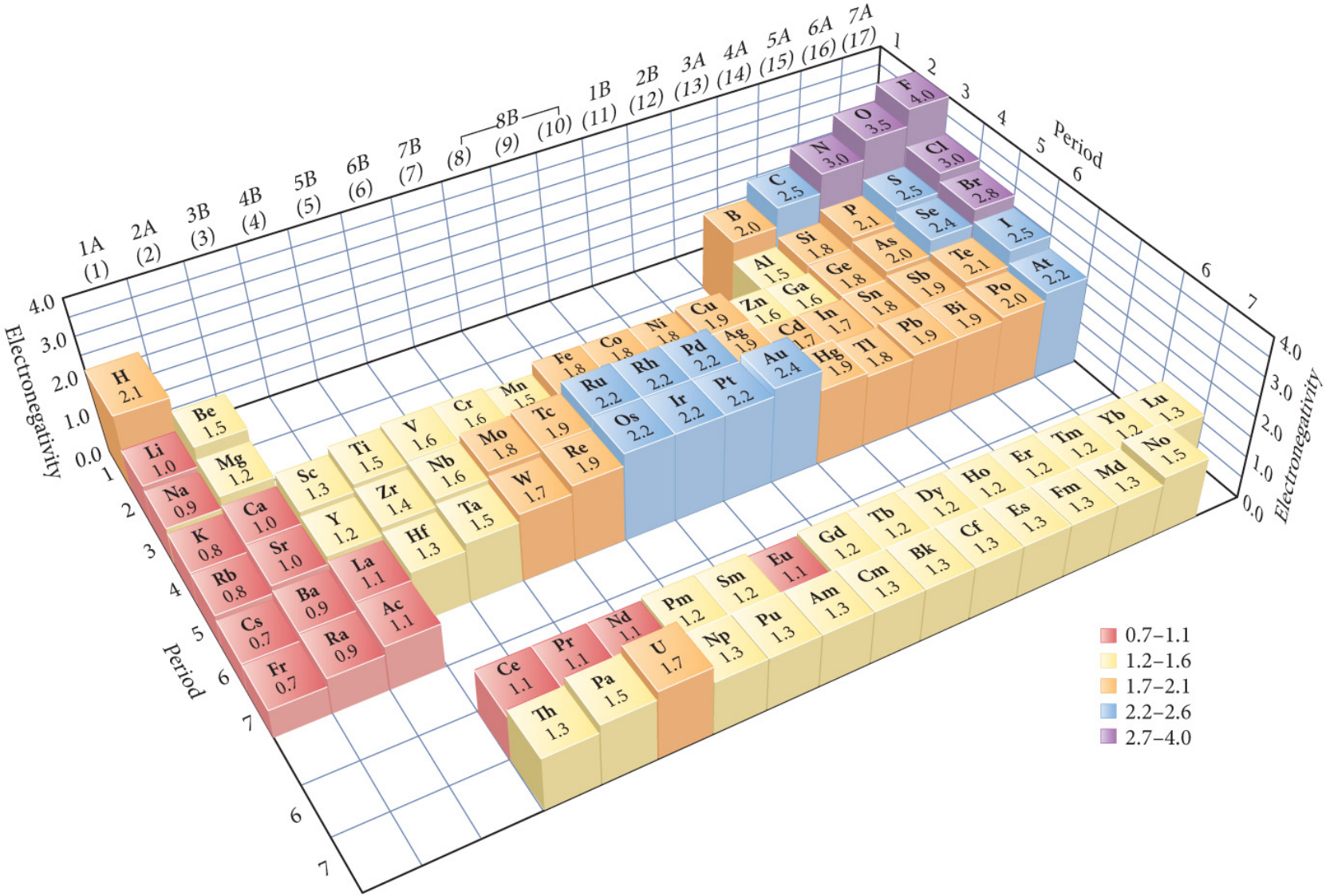
9.6: Electronegativity and Polar Bonds

Consider a covalent bond between two atoms: **X–Y**:

- The electrons in a covalent bond are not necessarily shared evenly between the two atoms.
- The **Electronegativity** of an element is a measure its ability to attract the shared electrons in a covalent bond.
- Electronegativity increases with increasing $Z_{\text{effective}}$, so electronegativity increases as you go up a group or left to right across a period (Figure 9.8).

Figure 9.8

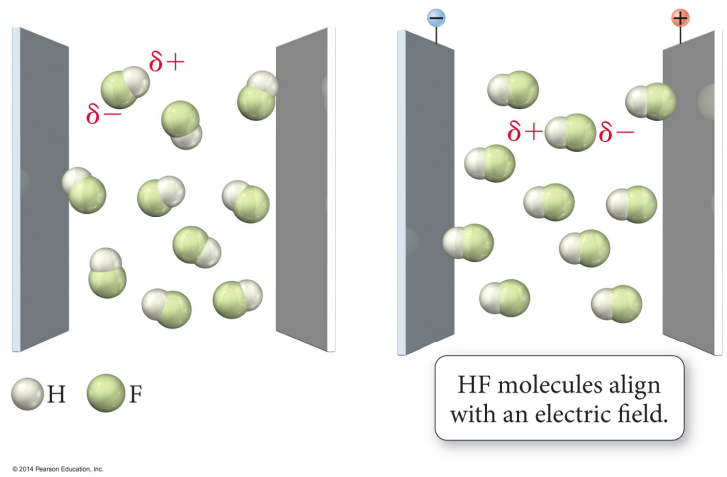
Trends in Electronegativity



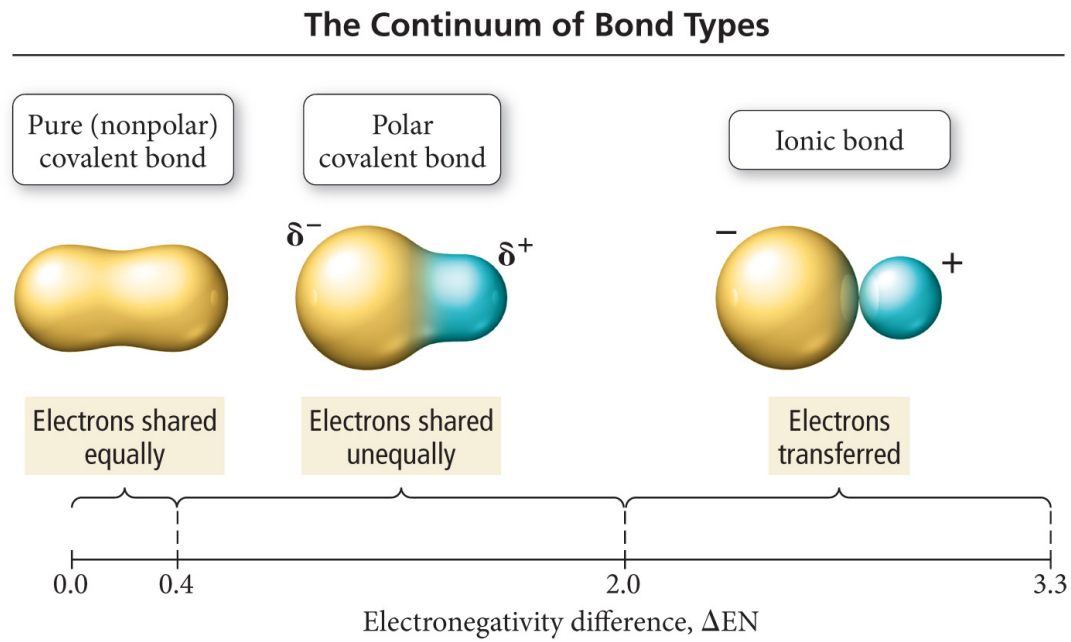
- Pauling scale of EN: Relative to F = 4.0

Electronegativity and Polar Bonds (contd.)

- The greater the electronegativity difference (ΔEN) in a bond, the **more uneven the electron sharing**.
- If a bond is between elements with different electronegativity (ΔEN), the electrons are shared unevenly. This is called a **polar covalent bond**.
- Polar bonds have both *covalent* and *ionic* character.



Electronegativity and Polar Bonds (contd.)



9.7: Lewis Structures of Molecules and Polyatomic Ions

Guidelines for drawing Lewis structures:

1) Arrange the atoms with the correct connectivity:
   - The least electronegative element is usually **central**
   - H and halogens are generally **terminal**
   - Symmetrical structures preferred
2) Calculate the <u>total number of valence electrons</u> in the molecule. A Lewis structure **<u>must</u>** contain this number of VE, in bonds or lone pairs.
   - For polyatomic ions, add or subtract electrons due to charge.
3) Join atoms with single bonds (uses 2 e^- /bond).
4) Add remaining VE as lone pairs to complete the octets of all the atoms (except H).
5) If there are not enough electrons to give each atom an octet, convert lone pairs to double or triple bonds.

Lewis Structures (contd.)

Try to follow “common bonding patterns”:

- **C** = 4 bonds & 0 lone pairs
- **N** = 3 bonds & 1 lone pair
- **O** = 2 bonds & 2 lone pairs
- **H** = 1 bond, no lone pairs (always)
- **halogens** = 1 bond & 3 lone pairs
- **B** = 3 bonds & 0 lone pairs

Lewis Structure Examples:

1) ammonia, NH_3

2) phosgene, $COCl_2$

Lewis Structure Examples (contd):

3) Hydrogen cyanide, HCN

4) Nitrate ion, NO_3^-