

Section 4.3: Electronegativity and Bond Polarity

Tutorial 1 Practice, page 220

$$\begin{aligned} 1. \text{ (a) } \Delta EN_{\text{F-C}} &= EN_{\text{F}} - EN_{\text{C}} \\ &= 4.0 - 2.5 \end{aligned}$$

$$\Delta EN_{\text{F-C}} = 1.5$$

Since $0.5 < \Delta EN < 1.7$, the C–F bond is polar covalent.

$$\begin{aligned} \text{(b) } \Delta EN_{\text{Cl-P}} &= EN_{\text{Cl}} - EN_{\text{P}} \\ &= 3.0 - 2.1 \end{aligned}$$

$$\Delta EN_{\text{Cl-P}} = 0.9$$

Since $0.5 < \Delta EN < 1.7$, the P–Cl bond is polar covalent.

$$\begin{aligned} \text{(c) } \Delta EN_{\text{Cl-Li}} &= EN_{\text{Cl}} - EN_{\text{Li}} \\ &= 3.0 - 1.0 \end{aligned}$$

$$\Delta EN_{\text{Cl-Li}} = 2.0$$

Since $\Delta EN > 1.7$, the Li–Cl bond is ionic.

$$\begin{aligned} 2. \text{ (a) } \Delta EN_{\text{H-B}} &= EN_{\text{H}} - EN_{\text{B}} \\ &= 2.1 - 2.0 \end{aligned}$$

$$\Delta EN_{\text{H-B}} = 0.1$$

$$\begin{aligned} \Delta EN_{\text{H-H}} &= EN_{\text{H}} - EN_{\text{H}} \\ &= 2.1 - 2.1 \end{aligned}$$

$$\Delta EN_{\text{H-H}} = 0.0$$

$$\begin{aligned} \Delta EN_{\text{C-H}} &= EN_{\text{C}} - EN_{\text{H}} \\ &= 2.5 - 2.1 \end{aligned}$$

$$\Delta EN_{\text{C-H}} = 0.4$$

$$\begin{aligned} \Delta EN_{\text{F-H}} &= EN_{\text{F}} - EN_{\text{H}} \\ &= 4.0 - 2.1 \end{aligned}$$

$$\Delta EN_{\text{F-H}} = 1.9$$

$$\begin{aligned} \Delta EN_{\text{O-H}} &= EN_{\text{O}} - EN_{\text{H}} \\ &= 3.5 - 2.1 \end{aligned}$$

$$\Delta EN_{\text{O-H}} = 1.4$$

$$\begin{aligned} \Delta EN_{\text{H-Na}} &= EN_{\text{H}} - EN_{\text{Na}} \\ &= 2.1 - 0.9 \end{aligned}$$

$$\Delta EN_{\text{H-Na}} = 1.2$$

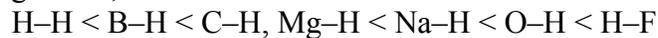
$$\begin{aligned} \Delta EN_{\text{H-Mg}} &= EN_{\text{H}} - EN_{\text{Mg}} \\ &= 2.1 - 1.2 \end{aligned}$$

$$\Delta EN_{\text{H-Mg}} = 0.9$$

$$\Delta EN_{\text{H-H}} = 0.0; \Delta EN_{\text{H-B}} = 0.1; \Delta EN_{\text{H-C}} = 0.4; \Delta EN_{\text{H-Mg}} = 0.9; \Delta EN_{\text{H-Na}} = 1.2;$$

$$\Delta EN_{\text{O-H}} = 1.4; \Delta EN_{\text{F-H}} = 1.9$$

The order of these molecules according to their predicted bond polarity, from least to greatest, is:



$$\begin{aligned} \text{(b)} \quad \Delta EN_{\text{Cl-Al}} &= EN_{\text{Cl}} - EN_{\text{Al}} \\ &= 3.0 - 1.5 \end{aligned}$$

$$\Delta EN_{\text{Cl-Al}} = 1.5$$

$$\begin{aligned} \Delta EN_{\text{Cl-P}} &= EN_{\text{Cl}} - EN_{\text{P}} \\ &= 3.0 - 2.1 \end{aligned}$$

$$\Delta EN_{\text{Cl-P}} = 0.9$$

$$\begin{aligned} \Delta EN_{\text{Cl-Li}} &= EN_{\text{Cl}} - EN_{\text{Li}} \\ &= 3.0 - 1.0 \end{aligned}$$

$$\Delta EN_{\text{Cl-Li}} = 2.0$$

$$\begin{aligned} \Delta EN_{\text{Cl-Cl}} &= EN_{\text{Cl}} - EN_{\text{Cl}} \\ &= 3.0 - 3.0 \end{aligned}$$

$$\Delta EN_{\text{Cl-Cl}} = 0.0$$

$$\begin{aligned} \Delta EN_{\text{Cl-I}} &= EN_{\text{Cl}} - EN_{\text{I}} \\ &= 3.0 - 2.5 \end{aligned}$$

$$\Delta EN_{\text{Cl-I}} = 0.5$$

$$\begin{aligned} \Delta EN_{\text{Cl-Rb}} &= EN_{\text{Cl}} - EN_{\text{Rb}} \\ &= 3.0 - 0.8 \end{aligned}$$

$$\Delta EN_{\text{Cl-Rb}} = 2.2$$

$$\Delta EN_{\text{Cl-Cl}} = 0.0; \Delta EN_{\text{I-Cl}} = 0.5; \Delta EN_{\text{P-Cl}} = 0.9; \Delta EN_{\text{Al-Cl}} = 1.5; \Delta EN_{\text{Li-Cl}} = 2.0;$$

$$\Delta EN_{\text{Rb-Cl}} = 2.2$$

The order of these molecules according to their predicted bond polarity, from least to greatest, is:



$$\begin{aligned} \text{(c) } \Delta EN_{\text{C-H}} &= EN_{\text{C}} - EN_{\text{H}} \\ &= 2.5 - 2.1 \end{aligned}$$

$$\Delta EN_{\text{C-H}} = 0.4$$

$$\begin{aligned} \Delta EN_{\text{C-F}} &= EN_{\text{F}} - EN_{\text{C}} \\ &= 4.0 - 2.5 \end{aligned}$$

$$\Delta EN_{\text{C-F}} = 1.5$$

$$\begin{aligned} \Delta EN_{\text{C-Cl}} &= EN_{\text{Cl}} - EN_{\text{C}} \\ &= 3.0 - 2.5 \end{aligned}$$

$$\Delta EN_{\text{C-Cl}} = 0.5$$

$$\begin{aligned} \Delta EN_{\text{C-O}} &= EN_{\text{O}} - EN_{\text{C}} \\ &= 3.5 - 2.5 \end{aligned}$$

$$\Delta EN_{\text{C-O}} = 1.0$$

$$\begin{aligned} \Delta EN_{\text{C-S}} &= EN_{\text{C}} - EN_{\text{S}} \\ &= 2.5 - 2.5 \end{aligned}$$

$$\Delta EN_{\text{C-S}} = 0.0$$

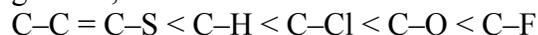
$$\begin{aligned} \Delta EN_{\text{C-C}} &= EN_{\text{C}} - EN_{\text{C}} \\ &= 2.5 - 2.5 \end{aligned}$$

$$\Delta EN_{\text{C-C}} = 0.0$$

$$\Delta EN_{\text{C-C}} = 0.0; \Delta EN_{\text{C-S}} = 0.0; \Delta EN_{\text{C-H}} = 0.4; \Delta EN_{\text{C-Cl}} = 0.5; \Delta EN_{\text{O-C}} = 1.0;$$

$$\Delta EN_{\text{C-F}} = 1.5$$

The order of these molecules according to their predicted bond polarity, from least to greatest, is:



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1. (a) $EN_{\text{C}} = 2.5$, $EN_{\text{N}} = 3.0$, $EN_{\text{O}} = 3.5$

The elements, listed in order of increasing electronegativity, are C, N, O.

(b) $EN_{\text{S}} = 2.5$, $EN_{\text{Se}} = 2.4$, $EN_{\text{Cl}} = 3.0$

The elements, listed in order of increasing electronegativity, are Se, S, Cl.

(c) $EN_{\text{Si}} = 1.8$, $EN_{\text{Ge}} = 1.8$, $EN_{\text{Sn}} = 1.8$

All 3 elements have the same electronegativity.

(d) $EN_{\text{Tl}} = 1.8$, $EN_{\text{S}} = 2.5$, $EN_{\text{Ge}} = 1.8$

The electronegativity of sulfur is higher than that of thallium and germanium, which both have the same electronegativity.

2. I could use the fact that electronegativity increases across periods (rows) and decreases across groups (columns). Bonds between atoms farther from each other in a period or group will have a higher difference in electronegativity (more of an ionic character), and bonds between atoms closer together in a period or group will have a lower electronegativity (more of a non-polar covalent character).

- 3. (a)** The C–O bond is polar covalent because the electronegativity of the elements increases rapidly as you move across Period 2, so the electronegativity difference between carbon and oxygen will be greater than 0.5.
- (b)** The F–I bond is polar covalent because the electronegativity of the halogens decreases rapidly as you move down Group 17, so the electronegativity difference between fluorine and iodine will be greater than 0.5.
- (c)** The Li–F bond is ionic because lithium has a low electronegativity and fluorine has the highest, so the electronegativity difference will be greater than 1.7.
- (d)** The Ge–Sn bond is non-polar covalent because these two elements are right beside one another in Group 14, in Periods 4 and 5, and after Period 2, the negativity of elements decreases or increases only gradually as you move down a group, so the electronegativity difference will be less than 0.5.
- (e)** The Al–Cl bond will be polar covalent because these two elements are in the same period, and electronegativity increases quickly as you move right from Group 13 (aluminum) to Group 17 (chlorine).
- 4. (a)** The Ge–F bond will be the most polar of C–F, Si–F, and Ge–F. C, Si, and Ge are all in Group 14, and electronegativity generally decreases as you move down a group. Ge has a lower electronegativity than C or Si, so the Ge–F bond will have the highest difference in electronegativity.
- (b)** The P–Cl bond will be the most polar of P–Cl and S–Cl. Sulfur and chlorine are in the same period, and electronegativity generally increases as you move across a period, so chlorine will have a lower electronegativity than sulfur, and thus the P–Cl bond will have a higher difference in electronegativity than the S–Cl bond will.
- (c)** The S–F bond will be the most polar of S–F, S–Cl, and S–Br. F, Cl, and Br are all halogens. Electronegativity will decrease as you move down the group. S is in the period between F and Br. The difference in electronegativity between S and Br is likely to be less than that between S and F, since F has the highest electronegativity of all the elements.
- (d)** The Ca–Cl bond will be the most polar of Be–Cl, Mg–Cl, and Ca–Cl. Be, Mg, and Ca are in Group 2; Ca is the lowest element in that group so it will have the lowest electronegativity of Be, Mg, and Ca. Thus, of Be, Mg, and Ca, Ca will have the highest electronegativity difference with Cl.
- 5.** The bonds K–Br, C–Br, and Br–Br differ as follows: K has a low electronegativity (EN = 0.8) and creates an ionic bond with the highly electronegative Br (EN = 2.8). The electronegativity of C (EN = 2.5) is closer to the electronegativity of Br, and C creates a polar covalent bond with Br. Since the elements in the Br–Br bond are identical, the electronegativity difference between the bromine atoms is zero, resulting in a non-polar covalent bond.
- 6. (a)** $O\delta^+$, $F\delta^-$
- (b)** non-polar
- (c)** $C\delta^+$, $Br\delta^-$
- (d)** $C\delta^+$, $O\delta^-$

7. (a) The bond polarity $\delta^+ \text{H}-\text{Br}\delta^-$ is correct.
 (b) The bond polarity is shown incorrectly; the correct bond polarity is $\delta^+ \text{I}-\text{Cl}\delta^-$.
 (c) The bond polarity $\delta^+ \text{Si}-\text{S}\delta^-$ is correct.
 (d) The bond polarity is shown incorrectly; the F-F bond is non-polar.
 (e) The bond polarity is shown incorrectly; the correct bond polarity is $\delta^+ \text{P}-\text{O}\delta^-$.

8. (a) $\Delta \text{EN}_{\text{F-C}} = \text{EN}_{\text{F}} - \text{EN}_{\text{C}}$
 $= 4.0 - 2.5$

$$\Delta \text{EN}_{\text{F-C}} = 1.5$$

The electronegativity difference between C and F is 1.5.

- (b) The C-F bond is polar covalent.
 (c) When you place molecules of CF_4 in an electric field, the molecules do not display any particular orientation because the positive and negative charges on the various atoms are uniformly distributed due to the molecular structure, so the molecule does not have a negative or a positive end.
9. (a) Since nitrogen has an electronegativity of 3.0, only atoms that satisfy $\text{EN} \geq 3.5$ will form a polar covalent bond with nitrogen where nitrogen has a partial positive charge. The only two atoms that meet this criterion are O and F.
 (b) Of O and F, F would result in the most polar bond with N
 (c) Since nitrogen has an electronegativity of 3.0, only atoms that satisfy $2.5 < \text{EN} < 3.5$ will form a non-polar covalent bond with nitrogen. Since $\text{EN}_{\text{Br}} = 2.8$ and $\text{EN}_{\text{Cl}} = 3.0$, both Br and Cl would result in a non-polar covalent bond when bonded to nitrogen. An N-N bond would also be non-polar covalent.

10. (a) $\Delta \text{EN}_{\text{Cl-Rb}} = \text{EN}_{\text{Cl}} - \text{EN}_{\text{Rb}}$
 $= 3.0 - 0.8$

$$\Delta \text{EN}_{\text{Cl-Rb}} = 2.2$$

Since $\Delta \text{EN} > 1.7$, the bond between Rb and Cl is ionic.

- (b) The bonding atoms, S and S, are identical, so they have the same electronegativity values. The bond between these two atoms is therefore non-polar covalent.

(c) $\Delta \text{EN}_{\text{F-C}} = \text{EN}_{\text{F}} - \text{EN}_{\text{C}}$
 $= 4.0 - 2.5$

$$\Delta \text{EN}_{\text{F-C}} = 1.5$$

Since $0.5 < \Delta \text{EN} < 1.7$, the bond between C and F is polar covalent.

(d) $\Delta \text{EN}_{\text{C-Ba}} = \text{EN}_{\text{C}} - \text{EN}_{\text{Ba}}$
 $= 2.5 - 0.9$

$$\Delta \text{EN}_{\text{C-Ba}} = 1.6$$

Since $0.5 < \Delta \text{EN} < 1.7$, the bond between Ba and C is polar covalent.

(e) $\Delta \text{EN}_{\text{Se-B}} = \text{EN}_{\text{Se}} - \text{EN}_{\text{B}}$
 $= 2.4 - 2.0$

$$\Delta \text{EN}_{\text{Se-B}} = 0.4$$

Since $\Delta \text{EN} < 0.5$, the bond between B and Se is non-polar covalent.

$$\begin{aligned} \text{(f) } \Delta EN_{\text{Br-Cs}} &= EN_{\text{Br}} - EN_{\text{Cs}} \\ &= 2.8 - 0.7 \end{aligned}$$

$$\Delta EN_{\text{Br-Cs}} = 2.1$$

Since $\Delta EN > 1.7$, the bond is ionic.