## Answers to Review for Chapter 4: Structure and Properties of Substances

### Chapter 4 review (p. 209 – 211): Q 2, 3, 6, 8, 9, 13 – 17, 19, 20, 24 – 27, 29 – 31

2. Inter-molecular forces in the hydrogen halides:

<table>
<thead>
<tr>
<th></th>
<th>HF</th>
<th>HCl</th>
<th>HBr</th>
<th>HI</th>
</tr>
</thead>
<tbody>
<tr>
<td>molecular shape</td>
<td>linear</td>
<td>linear</td>
<td>linear</td>
<td>linear</td>
</tr>
<tr>
<td>symmetrical</td>
<td>no</td>
<td>no</td>
<td>no</td>
<td>no</td>
</tr>
<tr>
<td>ΔEN</td>
<td>1.78</td>
<td>0.96</td>
<td>0.76</td>
<td>0.46</td>
</tr>
<tr>
<td>inter-molecular forces</td>
<td>LDFs dipole-dipole H-bonding</td>
<td>LDFs dipole-dipole</td>
<td>LDFs weak dip-dip</td>
<td>LDFs</td>
</tr>
</tbody>
</table>

HF has the highest boiling point because it has the strongest inter-molecular attraction due to the extreme polarity of the H – F bond and the ability to form hydrogen bonds between molecules.

3. CHBr₃

3. CHBr₃

6. What types of inter-molecular attraction must be broken to melt the following?

<table>
<thead>
<tr>
<th></th>
<th>NH₃</th>
<th>NaI</th>
<th>Fe</th>
<th>CH₄</th>
</tr>
</thead>
<tbody>
<tr>
<td>molecular shape</td>
<td>AX₃E₁</td>
<td>linear</td>
<td>not applicable</td>
<td>AX₄E₀ tetrahedral</td>
</tr>
<tr>
<td>symmetrical</td>
<td>no</td>
<td>no</td>
<td>not applicable</td>
<td>yes</td>
</tr>
<tr>
<td>ΔEN</td>
<td>0.84</td>
<td>1.73</td>
<td>0</td>
<td>0.35</td>
</tr>
<tr>
<td>type of bonding</td>
<td>polar covalent</td>
<td>ionic</td>
<td>metallic</td>
<td>non-polar covalent</td>
</tr>
<tr>
<td>inter-molecular forces to break so substance can melt</td>
<td>LDFs dipole-dipole H-bonding</td>
<td>ion to ion (LDFs)</td>
<td>metallic bonding</td>
<td>LDFs only</td>
</tr>
</tbody>
</table>

8. Hydrogen bonding will be stronger in H₂O than in NH₃ because the ΔEN for the O – H bond is 1.24 while the ΔEN for the N – H bond is only 0.84. While both molecules are asymmetrical and capable of hydrogen bonding, the more polar O – H bond will cause a stronger dipole and stronger hydrogen bonding.

9. O₂²⁻

9. O₂²⁻

13. 

13. 

<table>
<thead>
<tr>
<th></th>
<th>C₃H₈</th>
<th>C₂H₅OH</th>
<th>K</th>
<th>SiO₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>type of bonding</td>
<td>non-polar covalent</td>
<td>polar covalent</td>
<td>metallic</td>
<td>covalent network</td>
</tr>
<tr>
<td>inter-molecular forces to break so substance can melt</td>
<td>LDFs only, so low boiling point</td>
<td>dip-dips, H-bonding and LDFs, so fairly high bp</td>
<td>metallic bonding in 3-D, so quite high boiling point</td>
<td>covalent bonding in 3-D, very high boiling point</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th></th>
<th>C₃H₈</th>
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<td>type of bonding</td>
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<td>polar covalent</td>
<td>metallic</td>
<td>covalent network</td>
</tr>
<tr>
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<td>covalent bonding in 3-D, very high boiling point</td>
</tr>
</tbody>
</table>
Answers to questions on page 209 and 210:

14.

<table>
<thead>
<tr>
<th></th>
<th>CCl₄</th>
<th>CH₂Cl</th>
<th>CHCl₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>molecular shape</td>
<td>AX₄E₀ tetrahedral</td>
<td>AX₂E₀ tetrahedral</td>
<td>AX₄E₀ tetrahedral</td>
</tr>
<tr>
<td>symmetrical</td>
<td>yes</td>
<td>no</td>
<td>no</td>
</tr>
<tr>
<td>ΔEN</td>
<td>0.61</td>
<td>0.35 and 0.61</td>
<td>0.35 and 0.61</td>
</tr>
<tr>
<td>type of bonding</td>
<td>non-polar covalent</td>
<td>polar covalent</td>
<td>polar covalent</td>
</tr>
<tr>
<td>inter-molecular forces to break so substance can melt</td>
<td>LDFs only</td>
<td>LDFs slight dipole-dipole</td>
<td>LDFs slight dipole-dipole</td>
</tr>
</tbody>
</table>

15. Dipole-dipole attraction is the inter-molecular attraction between partially charged (δ- and δ+) regions on polar molecules. They are relatively strong forces, and the strength increases as ΔEN increases. Ionic bonds are the intra-molecular forces of attraction between fully charged ions (- and +) in an ionic compound and are very strong. (Technically, the inter-molecular attraction between ionic compounds in a crystal lattice are the same as an ionic bond.)

16.

<table>
<thead>
<tr>
<th></th>
<th>NOCl</th>
<th>AlF₆³⁻</th>
<th>XeO₃</th>
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<tbody>
<tr>
<td>Lewis diagram</td>
<td>[diagram]</td>
<td>[diagram]</td>
<td>[diagram]</td>
</tr>
<tr>
<td>molecular shape</td>
<td>AX₂E₁ bent (V-shaped)</td>
<td>AX₃E₀ octahedral</td>
<td>AX₃E₁ trigonal pyramidal</td>
</tr>
<tr>
<td>symmetrical</td>
<td>no</td>
<td>yes</td>
<td>no</td>
</tr>
</tbody>
</table>

17. a) A compound of chlorine (therefore more than one type of atom) that is diatomic (contains 2 atoms) and has only dispersion forces between the molecules when in the liquid state (so ΔEN must be less than 0.50). Either Cl-Br or Cl-I would be correct answers:

```
Cl − Br
```

b) A compound of chlorine (therefore more than one type of atom) that is diatomic (contains 2 atoms) and has dipole-dipole attraction between the molecules when in the liquid state (so ΔEN must be between 0.50 and 1.70). Either H-Cl or Cl-F would be correct answers (Cl-Br would also technically have a very slight dipole):

```
Cl − H
```

19. A molecule with polar bonds can be a non-polar molecule when the molecule is symmetrical in all planes, because then the dipoles act equally in opposite directions and cancel out.
Answers to questions on page 210 :

20. PF₂Cl₃ can be drawn so that is both polar and non-polar.

Polar  because asymmetrical  
Non-polar  because symmetrical

24. XeF₆ will NOT have an octahedral shape because there is a lone pair of electrons on the central Xe atom. This is because xenon has 8 valence electrons. Six of these electrons are found in the bonds with fluorine, leaving 2 electrons left over as a lone pair.

25. Based on ∆EN values, HF has more polar bonds, so it should have a higher boiling point than water. However, water actually has the higher boiling point because of its bent shape. Each water molecule can form hydrogen bonds with several other water molecules, making a strong network of hydrogen bonds. HF, because of its linear shape, can only form hydrogen bonds with a few surrounding molecules, so it has smaller inter-molecular forces and a lower boiling point than water.

26. SF₄ has the VSEPR notation AX₄E₁, so its shape is seesaw. Seesaw is an asymmetrical shape, so this molecule is polar. SiF₄ has the VSEPR notation AX₄E₀ so it is a tetrahedral shape. This is a symmetrical shape, so SIF₄ is a non-polar molecule.

27. NaI is an ionic compound, so ion to ion attractions (from the crystal lattice) must be overcome in order to melt samples of this compound. HI is a slightly polar covalent molecule (asymmetrical and ∆EN = 0.46), so there will be weak dipole-dipole attractions and LDFs that must be overcome to allow this substance to melt.

29. CO₂ is a linear molecule that is symmetrical, so it is non-polar which makes it a gas at room temperature. SiO₂ is a covalent network solid that is bonded in three dimensions, which makes it a solid at room temperature.
30. SeO$_3$ has VSEPR notation AX$_3$E$_0$ so while it contains polar bonds, it is trigonal planar in shape, symmetrical and non-polar. SeO$_3$ has VSEPR notation AX$_3$E$_1$, so it contains polar bonds. It is bent in shape, so it is asymmetrical which means it is a polar molecule.

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\begin{array}{c}
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\cdot \cdot \cdot \\
\ \\
\cdot \cdot \\
\end{array}
\]

31. N$_2$H$_4$ has non-polar covalent bonds between the two nitrogen atoms and polar covalent bonds between the nitrogen and hydrogen atoms. Because of the lone pairs of electrons on the nitrogen atoms, the molecule is asymmetrical, which makes it a polar so it will have dipole-dipole inter-molecular attractions, LDFs and hydrogen bonding. C$_2$H$_4$ has a non-polar double covalent bond between the two carbon atoms and non-polar covalent bonds between the carbon and hydrogen atoms. Because there are no lone pairs of electrons on the carbon atoms, the molecule is symmetrical, which makes it non-polar. It will have only LDFs between molecules.

\[
\begin{array}{c}
\cdot \cdot \\
\ \\
\cdot \cdot \\
\ H \ \\
\ \\
\cdot \cdot \cdot \\
N \ - \ N \ - \ H \ \\
\ \\
\ H \ \\
\end{array}
\]

\[
\begin{array}{c}
\cdot \cdot \\
\ \\
\cdot \cdot \\
\ H \ \\
\ \\
\cdot \cdot \\
\ C \ = \ C \ = \ H \ \\
\ \\
\ H \ \\
\end{array}
\]
Be able to apply what we have learned. For example, predict four physical properties of phosgene (COCl$_2$), an extremely toxic compound that has been used for chemical warfare.

Phosgene is a polar covalent compound (it is asymmetrical and has polar bonds) so it will have fairly strong dipole-dipole attractions in addition to LDFs. This means that phosgene will probably be a gas or liquid at SATP, have fairly high melting and boiling points, is soluble in water, has a distinctive odour and is a non-conductor of electricity both in its pure form and in solution.

Answers to questions on page 214:

<table>
<thead>
<tr>
<th>1. d</th>
<th>3. b</th>
<th>5. c (and also b?)</th>
<th>7. a</th>
<th>9. d</th>
</tr>
</thead>
<tbody>
<tr>
<td>2. b</td>
<td>4. d (and also c?)</td>
<td>6. e</td>
<td>8. c</td>
<td>10. b</td>
</tr>
</tbody>
</table>

13. $^{7}$N  
    [He] 2$s^2$ 2$p^3$

$^{15}$P  
    [Ne] 3$s^2$ 3$p^3$

$^{33}$As  
    [Ar] 4$s^2$ 3$d^{10}$ 4$p^3$. Because the d sub-level is full, can also write: $^{33}$As  
    [Ar] 3$d^{10}$ 4$s^2$ 4$p^3$

14. In order of increasing atomic radius (size):  
    Mg < Na < Ca < K

15. Molecules with bent shape have the VSEPR notation:  
    AX$_2$E$_1$ or AX$_2$E$_2$

16. Lone pairs of electrons take up more space than bonded pairs, so they tend to repel the electrons in bonds and push them closer together, decreasing the bond angles. For example, the bond angle in a tetrahedral molecule (AX$_4$E$_0$) is 109.5°, but the bond angle for a trigonal pyramidal molecule (AX$_3$E$_1$) with one lone pair is about 107°, and for a bent molecule (AX$_2$E$_2$) with two lone pairs is about 104.5°.

17. In order of increasing IE$_1$:
    a) Rb < K < Na < Li
    b) Li < Be < B < C

18. In period 4, chromium and copper both have different electron configurations than is predicted because these element promote a 4s electron to half-fill or fill the d-sublevel which gives a lower energy electron arrangement.

19. Lattice energy is related to the inter-molecular attraction between the ionic compounds in the crystal lattice. The larger the lattice energy, the higher the melting and boiling points and the harder the crystal.
21. The properties of the elements repeat in a periodic fashion because the properties depend on how many valence electrons are in the outer shell. All elements with one valence electron have similar properties, all elements that need only one valence electron to complete a full valence shell also have similar properties. Once the valence shell is full, the pattern starts again at the beginning of the next period.

22. The nuclear charge refers to the atomic number (Z), which is the number of protons in the nucleus of an atom. Hydrogen has a nuclear charge of 1+ and lithium has a nuclear charge of 3+.

The effective nuclear charge (Z_eff) refers to the actual "pull" exerted by the nucleus on the valence electrons. This is calculated by subtracting the number of shielding electrons from the nuclear charge. Both hydrogen and lithium have an effective nuclear charge of 1+.

23. The maximum number of electrons in an energy level is calculated using the equation 2n^2, where "n" is the principal quantum number of the energy level you are considering.

24. The uncertainly principle states that it is impossible to know BOTH where an electron is (its energy or position away from the nucleus) and where it is going (its path or trajectory). By measuring one of these parameters, we change the other. The implications of the uncertainty principle are that we can’t predict exactly where the electrons in an atom are going to be- there are no such thing as electron orbits. Instead, we talk about orbitals, which are three-dimensional regions where we are likely to find a particular electron 95% of the time.

25. Dipole-dipole attractions and LDFs are both types of inter-molecular attractions. Dip-dips are associated with permanently charged regions of molecules (full or partial charges) and are quite strong. Their strength increases with the ΔEN of the bonds in an asymmetrical molecule. LDFs are temporary very slightly charged regions on the atoms in molecules due to the random motion of electrons in the atom. LDFs are very weak, but increase as the size of atoms increases (because they have more electrons) and increase as the size and "straightness" of molecules increases.
29. For $n = 3$, $l = 2$, $m_l = -1$ and $m_s = + \frac{1}{2}$

The last electron is in the third energy level ($n = 3$), and the d orbital ($l = 2$). The possible values for $m_l$ are $-2, -1, 0, +1, +2$. If we follow the convention that the electrons fill in order, then $m_l$ tells us that this is the $d^2$ electron. Altogether, the last electron is $3d^2$, so the element must be titanium.

30. Given $\text{IE}_1 = 1.09$ easy to remove, so valence electron

$\text{IE}_2 = 2.35$ easy to remove, so valence electron

$\text{IE}_3 = 4.62$ easy to remove, so valence electron

$\text{IE}_4 = 6.22$ easy to remove, so valence electron

$\text{IE}_5 = 37.83$ hard to remove, so breaking stable octet

there are four valence electrons, so this element is in Group IVA (4) or IVB (14)

31. Of the ionic solids LiI and LiBr, I would expect LiBr to have the higher melting point because LiBr has a higher $\Delta EN$, so it will have stronger inter-molecular attraction.

32. Which element in each pair is less metallic?

a) Sb or As? Sb will have more metallic character because it is lower in the Group (has higher shielding effect), so it will lose its valence electrons more easily.

b) Si or P? Silicon will have more metallic character because it has weaker $Z_{eff}$, so it will lose its valence electrons more easily.

c) Be or Na? Na will have more metallic character because it has both more shielding effect and less $Z_{eff}$, so it will lose its valence electron much more easily than Be.

35. Zn, Cd and Hg all have the electron configuration ending in $d^{10}s^2$.

a) The valence electrons to participate in bonding are the $s^2$ electrons.

b) All three elements lose these $s^2$ electrons to form $2^+$ ions (mercury can also form $1^+$ ions)

c) Mercury is a liquid at SATP, so it has weaker metallic bonding than Zn or Cd

36. a) $\Delta EN$ N – H is $(3.04 – 2.20) = 0.84$ so this is a polar bond, N has $\delta$- charge

b) $\Delta EN$ F – N is $(3.98 – 3.04) = 0.94$ so this is a polar bond, F has $\delta$- charge

c) $\Delta EN$ I – Cl is $(3.16 – 2.66) = 0.50$ so this is defined as a non-polar bond, no partial charges (because it is right on the line, you could also say that this is a very slightly polar bond)

37. In order of increasing polarity:

a) Cl – Cl (no dipole) < Cl – Br (Cl is $\delta$-) < Cl – F (F is $\delta$-)

b) Si – Si (no dipole) < S – Cl (Cl is $\delta$-) < P – Cl (Cl is $\delta$-) < Si – Cl (Cl is $\delta$-)

38. Molecule $AY_3$ is polar, therefore it must be an asymmetrical molecule.

It could be: $AX_3E_1$ (trigonal pyramidal), $AX_3E_2$ (T-shaped) or $AX_3E_3$

It can’t be $AX_3E_0$ (trigonal planar) because this would be non-polar

40. $POF_3$

The molecule is asymmetrical and contains polar bonds, so it is a polar molecule and will have dipole.
41. CH₂Cl₂

The molecule is tetrahedral shape (AX₄E₀). It is asymmetrical and contains polar bonds, so it is a polar molecule and will have dipole.

42. Compare CℓO₄⁻¹ and OSCℓ₂. Use the “system” for bonding by the octet rule to determine the coordinate bonds:

CℓO₄⁻¹:
- have: 32 electrons
- need: 40 electrons to give all atoms a stable octet
- there are 8 electrons in bonds, or 4 bonds
- shape is tetrahedral
- there is no expanded valence because all of the bonds on chlorine are coordinate covalent

SOCl₂:
- have: 26 electrons
- need: 32 electrons to give all atoms a stable octet
- there are 6 electrons in bonds, or 6 bonds
- shape is trigonal pyramidal
- there is no expanded valence because all of the bonds on sulfur are regular or coordinate covalent bonds

43. Draw the first excited state for a sodium atom:

Ground state configuration: ¹¹Na 1s² 2s² 2p⁶ 3s¹
First excited state: ¹¹Na 1s² 2s² 2p⁶ 3p¹ (the 3s electron jumps to the next sub-level, the 3p)

45. Across the periodic table from left to right:
- metallic character decreases
- atomic size decreases and ionization energy increases

Down the periodic table from top to bottom:
- metallic character increases
- atomic size increases and ionization energy decreases

47. Non-polar covalent bonds have ΔEN of less than 0.50
Polar covalent bonds have ΔEN of 0.50 – 1.70
Ionic bonds have ΔEN of greater than 1.70

Because the ΔEN of the O – H bond in water is 1.24, this bond is classified as polar covalent.

48. Physical properties of solid metallic elements are that they are shiny, ductile, malleable, have relatively high melting and boiling points and are good conductors of heat and electricity.

Chemical properties of metallic elements are that they react with water to produce basic solutions, react with chlorine to form ionic compounds (which are electrolytes) and react with acids to produce hydrogen gas (from grade 11). All of these properties are due to fact that metals tend to lose electrons to form positive ions.
49. Successive ionization energies for any element always increase because you are pulling an electron away from an increasingly positively charged ion.

b) Any time you break a stable octet, the successive ionization energies will jump considerably. For example, sodium’s second ionization energy will be much higher than its first IE. Magnesium’s third ionization energy will be much higher than its second IE.

50. Hund’s rule states that electrons will half-fill the orbitals of the same sub-level before electrons double up in any of the orbitals. For a carbon atom, this would predict that the 1s orbital will be filled first, the 2s orbital will be filled next, and then one electron will go into each of the 2px and 2py orbitals rather than putting two electrons into the 2px orbital.

51. a) As ΔEN increases, the ionic character of the bond also increases.

b) Ionic character is zero for any bond in which two bonding atoms are identical, for example a Cl – Cl bond. We can infer that the electrons in this molecule are equally attracted to both chlorine atoms, so they are equally shared and the Cl – Cl bond is completely non-polar.

c) No bonds have 100% ionic character. This implies that although electrons in a bond may be very much more strongly attracted to one atom than the other, there is always at least a tiny amount of attraction to the other atom. Electrons are never 100% transferred to one atom.

d) There is no single exact “cut-off” point at which bonds can be defined as non-polar, polar or ionic. Instead, all bonds are on a continuum from completely non-polar (ΔEN=0) to increasingly polar. As bonds become more polar, their electrons are shifted more and more to towards the atom with higher electronegativity but they are always somewhat attracted to their original atom.

52. Electronegativity increases from left to right across the periodic table, while atomic radius decreases.

Electronegativity decreases from top to bottom down the periodic table, while atomic radius increases.

These trends make sense because as an atom has increasing electronegativity, it has a stronger pull on both its own electrons and on new electrons. This exerts a strong pull on new electrons as well as pulling its own electrons in tight to the nucleus, which decreases atomic size.

Similarly, going down a group, the atoms get bigger which means that the valence electrons are further from the nucleus. This means that the atom will exert a weaker pull on both new electrons and on its own valence electrons, so atomic size increases.

53. Heisenberg suggested the uncertainty principle, which implies that we never know for sure where an electron is going to be in an atom.

De Broglie suggested that electrons have wave-like properties, which also implies that we never know for sure where an electron is going to be.
Answers to Practice multiple choice:

<p>| | | | | | |</p>
<table>
<thead>
<tr>
<th></th>
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<tbody>
<tr>
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<td>b</td>
<td>37</td>
<td>d</td>
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<td>d</td>
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