CHEM1611 Worksheet 4: Intermolecular Forces and Introduction to Acids and Bases

Information

Intermolecular forces are the interactions *between* rather than *inside* molecules. They are responsible for many of the physical properties of substances, including their melting and boiling points.

In pure substances, there are 3 important intermolecular forces which may be present:

- *Dipole – dipole forces*. The dipole moment in a molecule will tend to align with those in its neighbours. This type of interaction is only possible if the molecule possesses a dipole.
- *Hydrogen bonds*. This is a particularly strong dipole – dipole interaction involving the interaction between the δ+ H atoms in very polar bonds and lone pairs on very electronegative atoms. Hydrogen bonding therefore requires the presence of both δ+ H atoms *and* electronegative atoms.
- *Dispersion forces*. These forces are present in *all* substances. At any moment in time, the electron density in a molecule or atom may not be symmetrical and this leads to a dipole moment. This momentary or *instantaneous* dipole moment *induces* matching dipoles in neighbouring molecules or atoms by polarizing their electron density.

  Dispersion forces increase with the number of electrons in a molecule or atom.

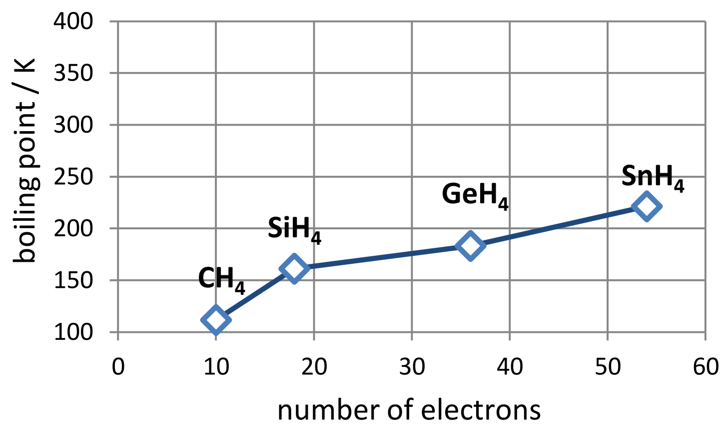
Model 1: Boiling Points Change Down a Group of the Periodic Table

Molecules are held in the liquid phase due to intermolecular forces so that boiling points are a good guide to their strength.

The figure opposite shows the boiling points of the Group 14 hydrides. All have the same shape but differ in the total number of electrons.

For example:

- C has 6 electrons and each H has 1 electron so CH_4 has $6 + 4 \times 1 = 10$.
- Sn has 50 electrons so SnH_4 has 54 electrons.



Critical thinking questions

1. What happens to the boiling point as the number of electrons increases?

2. What shape are the Group 14 hydrides?

3. Are dipole – dipole forces present in these molecules?

4. Is hydrogen bonding possible in these molecules?

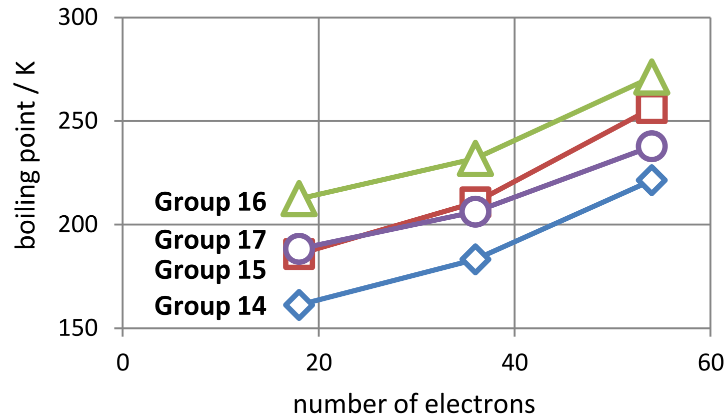
5. What intermolecular force is present in these molecules?

6. Explain why the boiling points vary in the way you described in answer to Q1.

Model 2: Boiling Points Change Across a Row of the Periodic Table

On the graph opposite, the boiling points for the other hydrides have been added:

- Group 14 SiH_4, GeH_4 and SnH_4
- Group 15 PH_3, AsH_3 and SbH_3
- Group 16 H_2S, H_2Se and H_2Te
- Group 17 HCl, HBr and HI



Critical thinking questions

1. Use a Periodic Table to confirm that SiH_4, PH_3, H_2S and HCl all have 18 electrons.

2. What happens to the boiling point as the number of electrons increases? Why?

3. What are the molecular shapes of PH_3, H_2S and HCl?

4. Do PH_3, H_2S and HCl have dipole moments?

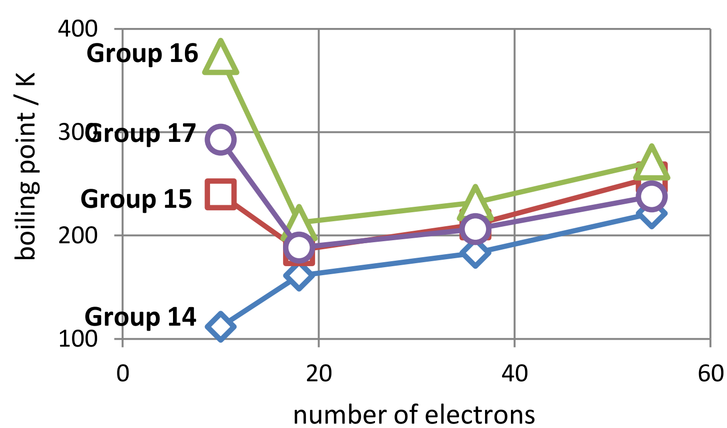
5. Why is the boiling point of SiH_4 *lower* than that of PH_3, H_2S and HCl?

6. Is the boiling point of SnH_4 (54 electrons) higher or lower than the boiling point of PH_3 (18 electrons)?

7. Explain your answer to Q6, making sure that it is consistent with your answers to Q2 and Q5.

Model 3: Anomalous Boiling Points of NH_3, H_2O and HF

The graph opposite adds the boiling points of CH_4, NH_3, H_2O and HF to Model 2. N, O and F are very electronegative and N-H, O-H and H-F bonds are very polar.



Critical thinking questions

1. How do the boiling points of the Group 14 hydrides change down the group?

Re-read your answers to Model 1.

2. How many δ+ H atoms are there on the most electronegative element in the molecules below?

(a) NH_3 (b) H_2O (c) HF

3. How many lone pairs are there on the most electronegative element in these molecules?

(a) NH_3 (b) H_2O (c) HF

4. Explain why the boiling points of NH_3, H_2O and HF (10 electrons) are *higher* than those of PH_3, H_2S and HCl (18 electrons) Refer to the **Information** if you are unsure.

5. Given your answer to Q4, suggest why the boiling point of NH_3 (10 electrons) is *lower* than that of SbH_3 (54 electrons).

6. Order the N-H, O-H and F-H bonds in terms of their polarity.

7. Predict the *relative* strength of the intermolecular forces between *two* NH_3 molecules, *two* H_2O and *two* HF molecules.

8. How many hydrogen bonds can each NH_3 molecule make on average in NH_3(l)? (*Hint*: re-read your answers to Q2 and Q3).

9. How many hydrogen bonds can each HF molecule make on average in HF(l)? (*Hint*: re-read your answers to Q2 and Q3).

10. How many hydrogen bonds can each H_2O molecule make on average in H_2O(l)? (*Hint*: re-read your answers to Q2 and Q3).

11. Use your answers to Q6 – Q10 to explain why the boiling points vary in the order $NH_3 < HF < H_2O$.

Model 4: pH

Water is able to act as both an acid and a base and it is possible for water to react with itself in an acid-base reaction called the *autoprotolysis* or *autoionization* of water:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

The equilibrium constant for this reaction $K_w = [H_3O^+(aq)][OH^-(aq)]$. At 25 °C, $K_w = 1.0 \times 10^{-14}$. Several definitions have proven to be useful:

$$pH = -\log_{10}[H_3O^+(aq)], \qquad pOH = -\log_{10}[OH^-(aq)] \qquad pK_w = -\log_{10}K_w$$

Critical thinking questions

1. During the course of a titration, a student measures the pH several times. What is $[H_3O^+(aq)]$ for each pH value below? (Actually calculate $[H_3O^+(aq)]$ – do *not* leave in the form 10^x).

pH	0.50	1.50	2.50	3.50	4.50	5.50	5.75
$[H_3O^+(aq)]$							

2. What is the effect of the number to the left and to the right of the decimal point in the pH on $[H_3O^+(aq)]$?

Model 5: Strong and Weak Acids

A **strong acid** is one that is essentially 100% dissociated in water: if 1.0 mole of the acid is added to enough water to make a 1.0 L solution, the solution will have $[H_3O^+(aq)] = 1.0$ M and will be pH = 0.00

A **weak acid** is one that is *significantly* less than 100% dissociated in water: if 1.0 mole of the acid is added to enough water to make a 1.0 L solution, the solution will have $[H_3O^+(aq)] < 1.0$ M and will be pH > 0.

When an acid HA is placed in water, $H_3O^+(aq)$ ions are produced according to the reaction below. The extent to which the reaction goes is given by the acid dissociation constant, K_a. The stronger the acid, the larger the value of K_a.

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq) \qquad K_a = \frac{[H_3O^+(aq)][A^-(aq)]}{[HA(aq)]}$$

Critical thinking questions

1. What are the *major* species present in a solution of a strong acid like HCl?

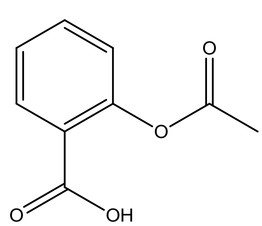
2. What are the *major* species present in a solution of a weak acid like CH_3COOH?

3. Under what pH conditions would $CH_3COO^-(aq)$ be the *dominant* species in a solution of CH_3COOH?

4. What are the *major* species present in a solution of a weak base like CH_3NH_2?

5. Under what pH conditions would $CH_3NH_3^+(aq)$ be the *dominant* species?

7. The extent of ionization of a drug helps determine how it is distributed in the body because ions are less likely to cross cell membranes than uncharged molecules. Are the two drugs below likely to be absorbed in (i) the acid environment of the stomach or (ii) the basic environment of the intestine?



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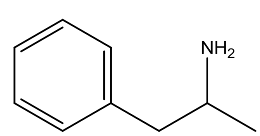

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