**4.4 Intermolecular forces**

**Learning objectives**

• Understand how intermolecular forces arise

• Understand how physical properties of covalent molecular substances depend on the intermolecular forces

• Predict relative boiling points of substances

**Understandings:**

 **Intermolecular forces include London (dispersion) forces, dipole-dipole forces and hydrogen bonding.**

**The relative strengths of these interactions are London (dispersion) forces < dipole-dipole forces < hydrogen bonds.**

**The nature of intermolecular forces**

**The force of attraction between molecules is called intermolecular force.**

There are various types of intermolecular forces. The main type between non-polar atoms/molecules is the **London (dispersion) force**. London forces are much weaker than covalent bonds.

London forces are present between all molecules in solid and liquid states.

**Example of bromine Br2**

Bromine is a liquid at room temperature. The bromine molecules are held together by weak London forces. Therefore, when bromine is heated to form a gas, the Br2 molecules (held together by covalent bonds) remain intact and it is the London forces that are overcome.

**Draw a diagram to illustrate Br2 in liquid state and gaseous state.**

**Van der Waals’ forces** is the collective name given to the forces between molecules and includes London (dispersion) forces, dipole−dipole interactions and dipole−induced dipole interactions.

**Intramolecular forces** are forces within a molecule.

**Intermolecular forces** are forces between molecules.

**Explain how London forces arise?**

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| **London forces are temporary (instantaneous) dipole - induced dipole interactions.**  |

The electrons in an atom are in constant motion, and at any one time they will not be symmetrically distributed about the nucleus. This results in a temporary (instantaneous) dipole in the atom, which will induce an opposite dipole in a neighbouring atom. These dipoles will attract each other so that there is an attractive force between atoms.

London forces get stronger as the number of electrons in molecules increases.

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| **In general, London forces get stronger as the relative molecular mass increases.**  |

 

The larger atoms in the molecule means that the outer electrons will be less strongly held, so the molecule is more polarisable, and therefore the induced dipoles will be larger.

 Variation in boiling point of alkanes with relative molecular mass

 

**Explain why butane has a higher boiling point than ethane?**

As the length (and so the relative molecular mass) of a hydrocarbon chain increases, so do the boiling points – butane (C4H10) has a higher boiling point than ethane (C2H6).

The forces between the molecules are stronger because there are more atoms, and therefore more electrons, present in butane than in ethane and also more points of contact between the chains.

**What is the effect of polar molecules on the nature and size of the intermolecular forces?**

Permanent dipole–permanent dipole interactions exist between molecules. These are shown in purple.

 

Because of the electronegativity difference between H and Cl, H–Cl molecules are polar.

London forces are present between the molecules in HCl(l) but, because of the polarity of the molecules, there are also other intermolecular forces present These are called **permanent dipole–permanent dipole** interactions, or usually just **dipole–dipole** attractions.

1. **Propane and ethanal have same molecular mass (44).Propane boiling point is -42C0 while ethanal boiling point is 21C0. Explain why the boiling point of ethanal is significantly higher? Draw the structure to illustrate your answer.**

 Propane C3H8 Ethanal CH3CHO

 CnH2n+2

Both molecules have similar molecular mass and therefore same LDF. However, ethanol has is polar and has dipole-dipole interactions which are stronger than the LDF.

1. **Explain why ICl (162.35) has a higher boiling point than Br2(159.80) despite their molecular masses being very close.**

**We cannot simply say polar molecules have higher boiling points without comparing masses**

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**Comparison of *cis-* and *trans*-1,2-dichloroethene. Explain why cis-1,2-dichloroethane has a higher boiling point than *trans*-1,2-dichloroethene**

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 These are isomers. What are isomers?

**Hydrogen bonding (intermolecular force)**

**Hydrogen bonding** influences many properties of substances – it is responsible for ice floating on water, it is the force between strands of DNA, it helps maintain the 3D structure of proteins, it is the reason that ethanol is soluble in water. It could be argued that hydrogen bonding is the reason why life on Earth exists as we know it!

**The origin of hydrogen bonding**

**Compare the boiling points of the hydrides of group 16 elements by use of the graph below.**

Boiling point/°C

120

100

80

60

40

20

0

–20

–40

–60

–80

H2O

20

40

60

80

100

120

140

Relative molecular mass

H2Te

H2S

H2Se

0

 Boiling points of group 16 hydrides.

**What is the origin of hydrogen bonding?**

**Draw a diagram to show hydrogen bonding in water**

Hydrogen bonding occurs between molecules when a very electronegative atom (N, O, F) is joined directly to a hydrogen atom in the molecule. The electronegative atom withdraws electron density from the hydrogen, polarising the bond such that there is a strong interaction between the δ+ hydrogen and the δ− atom (N, O, F) on the other molecule.

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| **The requirements for hydrogen bonding are that the H atom is attached to a very electronegative atom – N, O or F – which possesses at least one lone pair of electrons.**  |

**Diagrams showing hydrogen bonding in ammonia and HF**

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Hydrogen bonding can influence the solubility of substances in water, and often molecules that are able to hydrogen bond are soluble in water

 **Hydrogen bonding within molecules**

Consider the *cis* and *trans* forms of butenedioic acid shown in the Table below and explain why the *cis* form has a lower melting point.



Hydrogen bonding in ethanol Intramolecular hydrogen bonding in *cis*-but-2-ene-1,4-dioic acid 



 **Melting points and boiling points**

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| **Only intermolecular forces are broken when covalent molecular substances are melted or boiled – covalent bonds are not broken.**  |

The stronger the intermolecular forces, the more energy must be supplied to break them and the higher the boiling point.

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|  **London < permanent dipole–dipole < hydrogen bonding**  **WEAKEST STRONGEST**  |

London forces are present between all molecules and, in some substances, can provide a higher contribution to the intermolecular forces than dipole–dipole interactions – it is therefore important to consider substances with similar relative molecular masses.

**Arrange and explain sulfur, chlorine and argon in order of increasing boiling point and explain your order.**

**Compare the boiling points of propane (CH3CH2CH3), methoxymethane (CH3OCH3) and ethanol (CH3CH2OH)**

**Solubility**

It is often said, when referring to solubility, that ‘like dissolves like’. What this means is that: **generally a substance will dissolve in a solvent if the intermolecular forces in the solute and solvent are similar**

Whether or not a substance dissolves depends (in part) on how much energy is needed to overcome intermolecular forces in the solvent and solute and then how much energy is released, to pay back this energy, when intermolecular forces are formed between solvent and solute molecules in the solution.

**Substances that are able to participate in hydrogen bonding will generally be soluble in water, because they are able to hydrogen bond to the water.**

**Explain why Pentane is readily soluble in hexane but insoluble in water.**

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 **Explain why Ethanol is very soluble in water.**

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**Explain why longer chain alcohols become progressively less soluble in water.**

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**Describe dissolving ionic substances in water**

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Ion–dipole interactions **a** between water and sodium ions; **b** between water and chloride ions.

**Comparison of the physical properties of ionic and covalent molecular substances**

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Questions

1. a Describe the principles of the valence shell electron pair repulsion theory for predicting the shapes of molecules. [4]

b Predict the shapes and bond angles of the following molecules: [4]

 i PCl3

 ii CO2

c Explain why carbon dioxide is a non-polar molecule but sulfur dioxide is polar. [3]

d Draw a Lewis structure for carbon monoxide and explain whether it has a shorter or longer C–O bond length than carbon dioxide. [3]

1. Explain the following in terms of structure and bonding.
2. Sodium oxide has a high melting point and does not conduct electricity when solid but conducts electricity when molten. [4]
3. Sodium has a lower melting point than magnesium. [3]
4. Phosphine, PH3, has a lower boiling point than ammonia, NH3, and arsine, AsH3. [3]
5. Silicon dioxide has a much higher melting point than carbon dioxide. [3]