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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Structure 2.2
The Covalent Model

Booklet 3 · Understandings 2.2.8 – 2.2.12

Intermolecular forces, physical properties, chromatography, resonance and benzene

Complete revision notes & memory points · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Core material. Every student must know this. Understandings 2.2.8, 2.2.9 and 2.2.10 are common to both levels.

HL Additional higher level material. Understandings **2.2.11 (resonance structures)** and **2.2.12 (benzene)** are HL only. SL students can stop after section 4 and go straight to the key points and questions, skipping those marked HL.

Inside this booklet

1. What the syllabus actually asks of you
2. 2.2.8 — Intermolecular forces
3. 2.2.9 — Intermolecular forces and physical properties
4. 2.2.10 — Chromatography
5. HL 2.2.11 — Resonance structures
6. HL 2.2.12 — Benzene
7. Bringing it all together
8. Key points to memorise
9. Check yourself
10. Answers

1 What the syllabus actually asks of you

Booklets 1 and 2 dealt with the forces *inside* a molecule: the covalent bonds, the shape they produce and the polarity that results. This booklet turns to the much weaker forces *between* molecules. These intermolecular forces decide whether a covalent substance is a gas, a liquid or a solid at room temperature, whether it dissolves in water, and how it behaves on a chromatography plate. The HL sections then return to bonding itself, showing that some electrons are not confined to one bond at all but are spread over several atoms.

Understanding	What it says	What you must be able to do	Level
2.2.8 Intermolecular forces	The nature of the force between molecules is determined by their size and polarity. Intermolecular forces include London (dispersion), dipole–induced dipole, dipole–dipole and hydrogen bonding.	Deduce which intermolecular forces are present from the structural features of covalent molecules. Use "van der Waals forces" as the inclusive term for the first three.	SL+HL
2.2.9 Physical properties	Given comparable molar mass, the relative strengths of intermolecular forces are generally: London < dipole–dipole < hydrogen bonding.	Explain the volatility, electrical conductivity and solubility of covalent substances in terms of their structure.	SL+HL
2.2.10 Chromatography	Chromatography separates the components of a mixture by their relative attractions, through intermolecular forces, to a mobile phase and a stationary phase.	Explain, calculate and interpret retardation factor (R_F) values.	SL+HL
2.2.11 Resonance	Resonance structures occur when there is more than one possible position for a double bond in a molecule.	Deduce resonance structures of molecules and ions, using the term *delocalization*.	HL
2.2.12 Benzene	Benzene, C_6H_6, is an important example of a molecule that has resonance.	Discuss the structure of benzene from physical and chemical evidence.	HL

Linking questions and Nature of Science attached to 2.2.8 – 2.2.12

- **Nature of Science (Structure 1.1, 2.1, 2.3):** how do the terms "bonds" and "forces" compare? → section 2.7
- **Nature of Science:** how can advances in technology lead to changes in scientific definitions, such as the updated IUPAC definition of the hydrogen bond? → section 2.5
- **Structure 1.5:** to what extent can intermolecular forces explain the deviation of real gases from ideal behaviour? → section 2.8
- **Tool 1, Inquiry 2:** what experimental data demonstrate the physical properties of covalent substances? → section 3
- **Structure 3.2:** to what extent does a functional group determine the nature of the intermolecular forces? → section 3.2
- **Tool 1:** how can a mixture be separated using paper chromatography or thin layer chromatography (TLC)? → section 4
- HL **Structure 1.3:** why are oxygen and ozone dissociated by different wavelengths of light? → section 5.4
- HL **Reactivity 2.1, 2.2:** how does the resonance energy in benzene explain its relative unreactivity? → section 6.3
- HL **Reactivity 3.4:** what are the structural features of benzene that favour it undergoing electrophilic substitution reactions? → section 6.4

What the data booklet gives you — and what it does not

- **Given:** electronegativity values (Section 9), which tell you whether bonds are polar and whether hydrogen bonding is possible; bond lengths (Section 11) and bond enthalpies (Section 12) for single and double bonds, which you need for the resonance comparisons in sections 5 and 6; molar masses from the periodic table.
- **Not given:** boiling points, the strengths of intermolecular forces, R_F values, or the bond data for benzene and ozone. The order of strength, London < dipole–dipole < hydrogen bonding, and the three atoms (N, O, F) needed for hydrogen bonding must be memorised.

2 Structure 2.2.8 — Intermolecular forces SL + HL

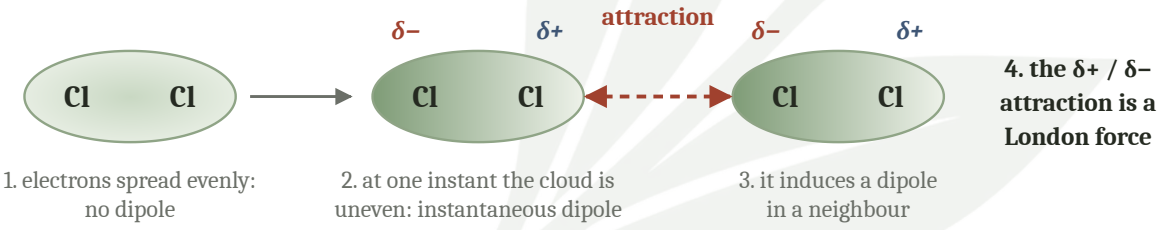
2.1 Intramolecular and intermolecular

The prefix *intra-* means "within" and *inter-* means "between". A covalent bond is an **intramolecular** force: it holds the atoms of one molecule together. An **intermolecular force** is an attraction between separate molecules. A sealed jar of chlorine contains countless Cl_2 molecules; the Cl–Cl bond within each molecule is strong (242 kJ mol^{-1}), but the attraction between neighbouring molecules is weak, which is why chlorine is a gas at room temperature.

All four intermolecular forces on the course are **electrostatic**: a region of partial negative charge on one molecule attracts a region of partial positive charge on another. They differ in where those partial charges come from — whether the dipoles are temporary or permanent — and that difference decides their strength.

2.2 London (dispersion) forces

Even in a non-polar molecule such as Cl_2, the electrons are constantly moving. At any instant, the electron cloud may happen to be denser at one end than the other. For that instant the molecule has a weak dipole, called an **instantaneous** (or temporary) **dipole**. Its δ− end repels the electrons of a neighbouring molecule, and its δ+ end attracts them, so the neighbour becomes polarized too: it acquires an **induced dipole**. The attraction between the instantaneous dipole and the induced dipole is a **London (dispersion) force**.



Cl_2 molecules: the instantaneous dipole changes constantly; the attraction is weak but always present

Figure 1 — An instantaneous dipole in one Cl_2 molecule induces a dipole in its neighbour.

- London forces act between **all** molecules, polar and non-polar alike, and between the atoms of the noble gases. They are the **only** intermolecular force between non-polar molecules.
- Individually they are the weakest intermolecular force, but their strength **increases with the number of electrons**, and so with molecular size and molar mass. A larger electron cloud is more easily distorted (more **polarizable**), so larger instantaneous dipoles form more often.

- They also increase with **surface area of contact**. Long, unbranched molecules touch along more of their length than compact, branched ones.

> **An exam habit worth forming**
>
> When asked to name the intermolecular forces present, **always include London forces**, whatever else is present. Water, for example, has London forces, dipole–dipole forces and hydrogen bonding. Leaving London forces out is a common way to lose a mark.

2.3 Dipole–dipole forces

A polar molecule such as HCl has a **permanent dipole**, because chlorine (3.2) is more electronegative than hydrogen (2.2). The δ+ end of one molecule attracts the δ− end of a neighbour. This **dipole–dipole attraction** acts in addition to the London forces that are always present. Because the dipoles are permanent rather than fleeting, dipole–dipole forces are stronger than London forces between molecules of similar size.

2.4 Dipole–induced dipole forces

In a mixture of polar and non-polar molecules, such as HCl with Cl_2, the permanent dipole of the polar molecule can distort the electron cloud of a non-polar neighbour, inducing a temporary dipole in it. The resulting attraction is a **dipole–induced dipole** force. It can only arise where polar and non-polar molecules are mixed, and it acts in addition to London forces.



permanent dipole – permanent dipole attraction

dipole–dipole: HCl with HCl



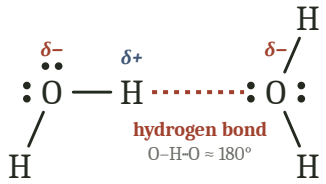
polar HCl — non-polar Cl_2, now polarized

dipole–induced dipole: HCl with Cl_2

> **Definition — van der Waals forces**
>
> **Van der Waals forces** is the inclusive term for **London (dispersion) forces, dipole–induced dipole forces and dipole–dipole forces** — every intermolecular attraction other than hydrogen bonding. Some older books use "van der Waals" to mean London forces alone; the IB uses it in the wider sense. Van der Waals forces can also act *within* a single large molecule when distant parts of it come close together, which matters for the folding of proteins.

2.5 Hydrogen bonding

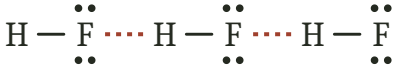
The strongest intermolecular force on the course arises when a hydrogen atom is covalently bonded to one of the three most electronegative small atoms: **nitrogen, oxygen or fluorine**. The large electronegativity difference pulls the bonding pair strongly away from hydrogen. Hydrogen has no inner electrons, so its nucleus is left almost unshielded, and this very small δ+ centre is strongly attracted to a **lone pair** on an N, O or F atom of a neighbouring molecule. That attraction is a **hydrogen bond**.



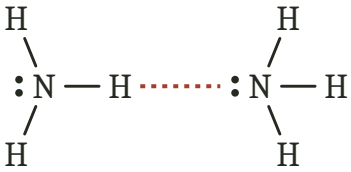
Definition — learn the IB wording

Hydrogen bonding occurs when a hydrogen atom, covalently bonded to an electronegative atom (N, O or F), has an attractive interaction with a lone pair on a neighbouring electronegative atom.

The three atoms H–X-Y lie in a straight line, because that brings the H closest to the lone pair. In a diagram, draw the hydrogen bond as a dashed or dotted line from the H to a lone pair on the other atom, and show the lone pair.



hydrogen fluoride: zigzag chains in the liquid



ammonia

A hydrogen bond needs **both** halves: an H attached to N, O or F on one molecule, and a lone pair on N, O or F on the other. Two consequences are often examined:

- Methane, CH_4, and other hydrocarbons cannot hydrogen bond: C (2.6) is not electronegative enough, and carbon has no lone pair.
- Propanone, CH_3COCH_3, and methoxymethane, CH_3OCH_3, cannot hydrogen bond *with each other*, because they have no H on O. They *can* accept a hydrogen bond from water, through the lone pairs on their oxygen atom, which is why they dissolve in water.

Nature of Science — how technology changes a definition

For much of the twentieth century the hydrogen bond was defined narrowly, as an attraction between H bonded to N, O or F and a neighbouring N, O or F. As infrared and NMR spectroscopy and X-ray and neutron diffraction became more sensitive, chemists detected weaker but genuine hydrogen-bond-like interactions involving other atoms. In 2011 IUPAC published a broader definition based on a set of experimental criteria — for example, a lengthening of the X–H bond and a change in its infrared absorption — that must be met before an interaction is called a hydrogen bond. Better instruments produced new evidence, and the definition was revised to fit it. For the IB exam, use the N, O and F definition above.

2.6 Deducing the forces present

Every molecule has London forces. → **Is the molecule polar?** (Booklet 2: polar bonds that do not cancel.) *Yes* → dipole–dipole forces as well. → **Is there an H on N, O or F**, and a lone pair on N, O or F? *Yes* → hydrogen bonding as well. → **A mixture** of polar and non-polar molecules? *Yes* → dipole–induced dipole forces too.

Intermolecular force		Non-polar molecules	Polar molecules	Molecules with H–N, H–O or H–F	Relative strength
London (dispersion)	van der Waals forces	✓	✓	✓	weakest for molecules of similar size
Dipole–induced dipole		only in a mixture of polar and non-polar molecules			weak
Dipole–dipole			✓	✓	stronger
Hydrogen bonding				✓	strongest

Substance	Polar?	H on N, O or F?	Intermolecular forces present
CO_2(s) (dry ice)	no	no	London only
N_2(l), I_2(s), CH_4	no	no	London only
HCl, CH_3Cl, CH_3OCH_3, CH_3CHO	yes	no	London + dipole–dipole
H_2O, NH_3, HF, CH_3OH	yes	yes	London + dipole–dipole + hydrogen bonding
CH_3OCH_3(l) mixed with hexane	one polar, one not	no	London + dipole–dipole (between CH_3OCH_3) + dipole–induced dipole

2.7 Bonds and forces: how do the terms compare?

Covalent bonds, ionic bonds, metallic bonds and every intermolecular force are all **electrostatic attractions** between opposite charges. The difference is one of **strength**, and the size of that difference is what justifies the two words. Breaking an O–H covalent bond needs 463 kJ mol^{-1}, but a hydrogen bond between water molecules is typically around 20 kJ mol^{-1} — a small fraction of a covalent bond — and London forces between small molecules are weaker still. Chemists keep the word **bond** for the strong attractions that hold atoms or ions together, and use **force** for the weak attractions between molecules. "Hydrogen bond" is the historical exception: it is an intermolecular force, despite its name.

2.8 Intermolecular forces and real gases

This is the Structure 1.5 linking question. The ideal gas model assumes there are **no forces** between gas particles. In a real gas there are, and they matter most at **low temperature** (particles move slowly, so attractions have time to act) and **high pressure** (particles are close together). Attractions pull particles together, reducing the pressure they exert below the ideal value. Gases with weak intermolecular forces, such as He and H_2, are the most nearly ideal; those with strong forces, such as NH_3 and H_2O, deviate most. Intermolecular forces explain a large part of the deviation — the rest comes from the volume of the particles themselves, which the ideal model also ignores.

Examiner traps in 2.2.8

- Leaving out London forces when listing the forces in a polar substance.
- Claiming hydrogen bonding in any molecule that contains hydrogen. The H must be bonded to **N, O or F**; C–H does not count.
- Drawing a hydrogen bond to the wrong atom, or without the lone pair on the acceptor.
- Calling a hydrogen bond a covalent bond, or saying it is "a bond between H and O within a molecule".
- Explaining London forces with "the molecule is polar". London forces arise from **instantaneous** dipoles, which every molecule has.

3 Structure 2.2.9 — Intermolecular forces and physical properties SL + HL

3.1 Melting and boiling: overcoming the forces between molecules

When a simple molecular substance melts or boils, the molecules are separated from one another, so it is the **intermolecular forces** that are overcome. The covalent bonds inside the molecules are not broken. Boiling water produces steam, which is still H_2O molecules, not separate hydrogen and oxygen atoms. Because intermolecular forces are weak, simple molecular substances have low melting and boiling points, and many are liquids or gases at room temperature. Covalent network structures are the contrast: melting them requires breaking strong covalent bonds, so diamond remains solid until roughly 4000 °C.

The rule — and the condition attached to it

Given **comparable molar mass**, the strength of intermolecular forces, and so the boiling point, generally increases in the order **London < dipole–dipole < hydrogen bonding**. The condition matters. A large non-polar molecule can have stronger London forces than the dipole–dipole forces of a small polar one, so always compare substances of similar molar mass before arguing from the type of force.

3.2 The effect of each factor, one at a time

(a) Number of electrons — London forces. Down group 17 and along the alkanes, the molecules are non-polar, so the only change is size. More electrons mean stronger London forces and higher boiling points:

Halogen	M / g mol^{-1}	Boiling point / °C	State at 25 °C
F_2	38	−188	gas
Cl_2	71	−34	gas
Br_2	160	59	liquid
I_2	254	184	solid

Alkane	M / g mol^{-1}	Boiling point / °C
CH_4	16	−162
C_2H_6	30	−89
C_3H_8	44	−42
C_4H_{10}	58	−1

(b) Shape — London forces. Pentane and its two branched isomers all have the formula C_5H_{12} and the same number of electrons, yet their boiling points differ: pentane 36 °C, 2-methylbutane 28 °C, 2,2-dimethylpropane 10 °C. Branching makes a molecule more compact, so it has less surface area in contact with its neighbours and weaker London forces.

(c) Type of force, at comparable molar mass. These three compounds have almost the same molar mass, so their London forces are similar. The differences in boiling point come from the additional forces:

Compound	M / g mol^{-1}	Polar?	H on N, O or F?	Intermolecular forces	Boiling point / °C
propane, $CH_3CH_2CH_3$	44	no	no	London only	−42
methoxymethane, CH_3OCH_3	46	yes	no	London + dipole–dipole	−24
ethanal, CH_3CHO	44	yes (C=O)	no	London + dipole–dipole	20
ethanol, CH_3CH_2OH	46	yes	**yes**	London + dipole–dipole + hydrogen bonding	**78**

Methoxymethane and ethanol are isomers (both C_2H_6O), so the 100 °C difference between them is entirely due to the O–H group. This is the Structure 3.2 linking question in action: the **functional group** largely decides which intermolecular forces are present — hydroxyl groups (alcohols) and carboxyl groups give hydrogen bonding, carbonyl groups give dipole–dipole forces — although the size and shape of the rest of the molecule matter too.

3.3 The hydrides of groups 14 to 17

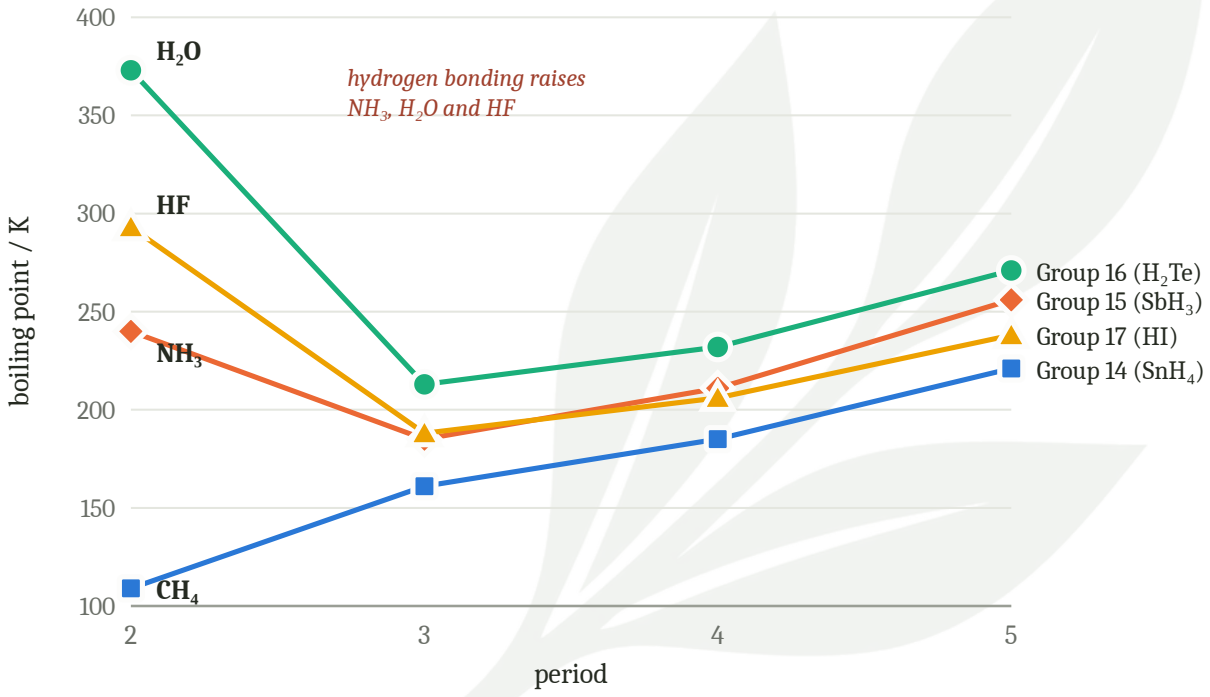


Figure 2 — Boiling points of the hydrides of groups 14–17. Each group is labelled at the end of its line; period 2 hydrides are labelled on the left.

In every group, the boiling point rises from period 3 to period 5 as the number of electrons, and so the strength of London forces, increases. Group 14 follows this trend all the way: CH_4 has the lowest boiling point because its molecules are non-polar and small.

NH_3, H_2O and HF are the anomalies. Their boiling points are far higher than the trend predicts, because their molecules form **hydrogen bonds**. Without hydrogen bonding, water would be a gas at room temperature, boiling at around −80 °C.

3.4 Water: the special case

Each water molecule has two H atoms and two lone pairs, so it can take part in up to **four** hydrogen bonds: two through its own hydrogen atoms and two through its lone pairs. In ice, every molecule is hydrogen bonded to four others in a tetrahedral arrangement. This holds the molecules in a rigid, open lattice with a lot of empty space, so **ice is less dense than liquid water** (about 0.92 compared with 1.00 g cm^{-3}). For almost every other substance, the solid is denser than the liquid. Because ice floats, ponds freeze from the top down and aquatic life survives beneath. The expansion of water on freezing also bursts pipes and splits rocks. Hydrogen bonds within large molecules matter too: they hold together the two strands of DNA and help to fix the shapes of proteins.

3.5 Volatility

Volatility is the tendency of a substance to vaporize. The weaker the intermolecular forces, the more volatile the substance and the lower its boiling point. Butane (C_4H_{10}, M = 58) is far more volatile than propan-1-ol (C_3H_7OH, M = 60), because the alcohol molecules are held by hydrogen bonds. Ionic compounds and covalent networks have very low volatility, because strong ionic or covalent bonds must be broken.

3.6 Solubility: "like dissolves like"

A substance dissolves when the attractions between solute and solvent particles can replace the attractions within the pure solute and the pure solvent.

- **Non-polar solutes dissolve in non-polar solvents.** Bromine and iodine dissolve in hexane or paraffin oil, because London forces form between solute and solvent.
- **Polar solutes dissolve in polar solvents such as water**, through dipole–dipole interactions and hydrogen bonds. Ethanol, glucose (with five O–H groups) and HCl are all very soluble in water.
- **Non-polar substances do not dissolve in water.** Water molecules are strongly hydrogen bonded to each other, and a non-polar molecule cannot form hydrogen bonds to replace them. Oil, wax, N_2 and hydrocarbons have very low solubility in water. Polar substances such as sugar are likewise insoluble in non-polar solvents such as cooking oil.
- **Chain length matters.** Ethanol mixes with water in all proportions, but heptan-1-ol, $C_7H_{15}OH$, is only slightly soluble. Its single O–H group can still hydrogen bond to water, but the long non-polar hydrocarbon chain cannot.
- **Covalent networks** such as diamond and SiO_2 are insoluble in every solvent: no solvent interaction releases enough energy to break their covalent bonds.

3.7 Electrical conductivity

Electric current needs mobile charged particles. Simple covalent substances contain neither ions nor delocalized electrons, so they **do not conduct** in any state. The exception is a polar molecule that **ionizes** in water: HCl(g) dissolves to form H^+(aq) and Cl^-(aq), so hydrochloric acid conducts. Among covalent networks, graphite and graphene conduct (delocalized electrons), silicon is a semiconductor, and diamond and SiO_2 are insulators (Booklet 2).

3.8 Summary of physical properties

	Ionic compounds	Polar covalent (molecular)	Non-polar covalent (molecular)	Covalent network
Volatility	low	higher	highest	low
Melting / boiling point	high	low to moderate	low	very high
Solubility in water	usually soluble	soluble; increases with polarity and hydrogen bonding	insoluble	insoluble
Solubility in hexane	insoluble	increases as polarity decreases	soluble	insoluble
Electrical conductivity	conducts only when molten or in aqueous solution	does not conduct (unless it ionizes in water)	does not conduct	does not conduct, except graphite and graphene; Si is a semiconductor

These are general trends, not universal rules. They are also exactly the experimental data the Tool 1 / Inquiry 2 linking question asks about: measuring melting and boiling points, testing solubility in water and in a non-polar solvent, and testing conductivity of the solid, the liquid and a solution are how chemists identify the structure type of an unknown substance.

Worked example — put CH_3CHO, CH_3CH_2OH and $CH_3CH_2CH_3$ in order of increasing boiling point, and explain your order.

1. Check molar masses: 44, 46 and 44 g mol^{-1}. These are almost identical, so the London forces will be similar, and the comparison can be made on the type of force.

2. Identify the additional forces: propane is non-polar (London only); ethanal is polar because of the C=O group (London + dipole–dipole); ethanol has an O–H group (London + dipole–dipole + hydrogen bonding).

3. Order: $CH_3CH_2CH_3$ < CH_3CHO < CH_3CH_2OH. Stronger intermolecular forces need more energy to overcome, so the boiling point is higher.

Model answer: "Explain why the boiling point of water is much higher than that of hydrogen sulfide." (3 marks)

Water molecules form hydrogen bonds with each other, because hydrogen is bonded to oxygen, which is highly electronegative and has lone pairs **[1]**. Hydrogen sulfide molecules have only London and dipole–dipole forces, because sulfur is not electronegative enough for hydrogen bonding **[1]**. Hydrogen bonds are stronger, so more energy is needed to separate water molecules **[1]**.

Examiner traps in 2.2.9

- Writing that boiling a covalent substance "breaks the covalent bonds". It overcomes the **intermolecular forces**.
- Comparing the type of force between molecules of very different molar mass. Check the molar masses first.
- Explaining a higher boiling point by "stronger bonds" without saying which forces, between which particles.
- Saying hydrocarbons are insoluble in water "because they are covalent". Ethanol and glucose are covalent too. The reason is that hydrocarbons are **non-polar** and cannot hydrogen bond with water.
- Saying a solution of a covalent substance conducts because of "free electrons". Where it conducts, it is because the substance has formed **ions**.

4 Structure 2.2.10 — Intermolecular forces and chromatography SL + HL

4.1 The principle

Chromatography separates the components of a mixture using two phases. The **stationary phase** does not move. The **mobile phase**, a solvent (or, in gas chromatography, a gas), moves through or over it. Each component is attracted to both phases through intermolecular forces, and spends part of its time dissolved in the mobile phase and part held by the stationary phase. A component that is more strongly attracted to the mobile phase travels further; one that is more strongly attracted to the stationary phase is held back. The components therefore separate.

4.2 Paper chromatography and thin layer chromatography (TLC)

	Paper chromatography	Thin layer chromatography (TLC)
Stationary phase	water held in the cellulose fibres of the paper (paper contains about 10% water, bound to the –OH groups of cellulose by hydrogen bonds)	a thin, even layer of silica (SiO_2) or alumina (Al_2O_3) on a glass, metal or plastic sheet; their surfaces carry –OH groups, so they are polar and can hydrogen bond
Mobile phase	a solvent, which rises through the stationary phase by capillary action. Several solvents may have to be tried before the components separate well.	
Advantages	simple and cheap	faster, gives sharper spots and better separation; the plate can be chosen to suit the mixture

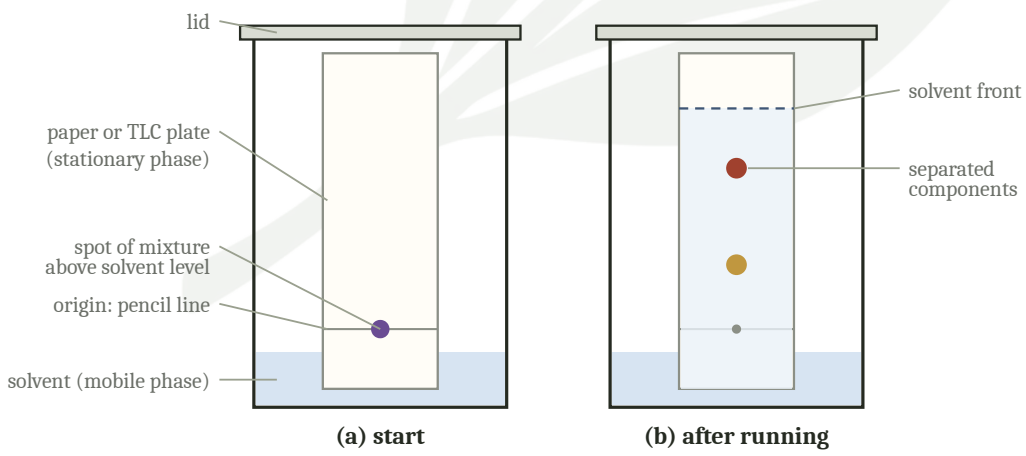


Figure 3 — Running a chromatogram.

1. Draw the **origin** line in **pencil** near the bottom of the paper or plate. Ink would dissolve in the solvent and run.
2. Apply a small, concentrated **spot** of the mixture on the origin, and spots of known reference substances beside it if you want to identify the components.
3. Stand the paper in a tank containing a shallow layer of solvent, with the origin **above the solvent level** (otherwise the sample dissolves into the solvent pool), and put a **lid** on so the tank stays saturated with solvent vapour.
4. When the solvent is close to the top, remove the paper and immediately mark the **solvent front** in pencil. Let it dry. The result is a **chromatogram**.

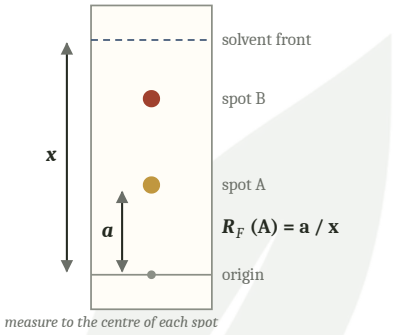
Colourless components have to be made visible after the run, for example with ultraviolet light or a locating agent such as ninhydrin for amino acids. The guide states that locating agents themselves are not required knowledge, and that the operational details of gas chromatography and HPLC will not be assessed.

4.3 The retardation factor, R_F

Each component is described by its **retardation factor**:

$$R_F = \frac{\text{distance travelled by the component}}{\text{distance travelled by the solvent}}$$

- Measure both distances **from the origin**, not from the bottom of the paper. Measure to the **centre** of each spot and to the solvent front, not to the top of the paper.
- R_F has **no units** and always lies between 0 and 1.
- Under identical conditions (same stationary phase, solvent and temperature), a substance always has the same R_F, so a component can be identified by comparing its R_F with a reference value or a reference spot run alongside it.

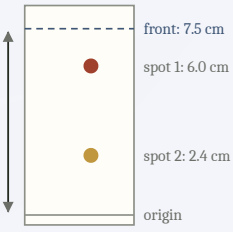


Worked example — calculating R_F

A mixture of two substances was separated by paper chromatography. The solvent front travelled 7.5 cm from the origin; spot 1 travelled 6.0 cm and spot 2 travelled 2.4 cm. Calculate both R_F values.

Spot 1: R_F = 6.0 / 7.5 = **0.80** Spot 2: R_F = 2.4 / 7.5 = **0.32**

Give R_F to two significant figures, the same precision as the measurements, and with no units.



4.4 Interpreting R_F values with intermolecular forces

The R_F value tells you which phase the component prefers. A **high** R_F means the component is strongly attracted to the **mobile phase** and weakly held by the stationary phase. A **low** R_F means the reverse. In TLC on polar silica with a **non-polar** solvent such as hexane:

- a **polar** component, especially one that can hydrogen bond (an –OH or –COOH group), is strongly attracted to the polar silica and travels only a short way: **low R_F**;
- a **non-polar** component is attracted more to the non-polar solvent than to the silica and travels further: **high R_F**.

Changing to a more polar solvent increases the R_F values of polar components, because the solvent now competes better with the stationary phase for them.

Model answer: "In TLC on silica with a non-polar solvent, X has R_F = 0.30 and Y has R_F = 0.90. Outline why Y travels further." (2 marks)

Y is less polar than X, so it is less strongly attracted to the polar silica stationary phase **[1]** and more strongly attracted to the non-polar mobile phase, so it spends more time moving with the solvent **[1]**.

Examiner traps in 2.2.10

- Measuring from the bottom of the paper, or to the top of the paper instead of the solvent front.
- Giving R_F with units, or a value greater than 1 (you have divided the wrong way round).
- Drawing the origin in ink, or placing it below the solvent level.
- Saying a component "travels further because it is lighter". The separation depends on the relative **attractions** (intermolecular forces) to the two phases, not on mass.
- Comparing R_F values measured with different solvents or plates.

5 Structure 2.2.11 — Resonance structures HL

5.1 When one Lewis formula is not enough

Ozone, O_3, has 18 valence electrons. Its Lewis formula has one O=O double bond and one O–O single bond, with the double bond on either side. Both arrangements are equally valid:



If either structure were the real one, ozone would contain two different oxygen–oxygen bonds, one long and weak and one short and strong. Experiment shows something else: the two bonds in ozone are **identical**, and intermediate in length and strength between a single and a double bond.

	O–O single bond	O=O double bond	Oxygen–oxygen bonds in O_3
Bond length / pm	148	121	128 (both)
Bond enthalpy / kJ mol^{-1}	144	498	≈ 364 (both)

Single- and double-bond values from the data booklet (Sections 11 and 12); ozone values are experimental.

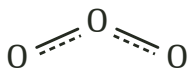
5.2 Delocalization and the resonance hybrid

The explanation is that the electrons of the second bonding pair are not confined between one particular pair of atoms. They are spread over all three oxygen atoms. Electrons shared by more than two atoms are called **delocalized**. Spreading electrons over a larger region lowers their energy, so delocalization makes the molecule or ion **more stable**.

Definitions to learn

- **Resonance structures** — two or more Lewis formulas that can be drawn for the same species, differing only in the position of a double bond (and of lone pairs). They are linked by a double-headed arrow, ↔.
- **Resonance hybrid** — the true structure, a blend of all the resonance structures, in which the electrons are delocalized. It is drawn with dashed lines showing where the delocalized electrons are spread.
- **Delocalization** — the sharing of a pair of electrons between three or more atoms.

The molecule does **not** switch rapidly between the resonance structures. Neither structure exists on its own: the real molecule is a single, unchanging hybrid. A useful analogy is a mule, the offspring of a horse and a donkey. Describing it as "part horse, part donkey" does not mean it is a horse one moment and a donkey the next. On a resonance hybrid, lone pairs are not drawn, because some of them take part in the delocalization and have no fixed position.



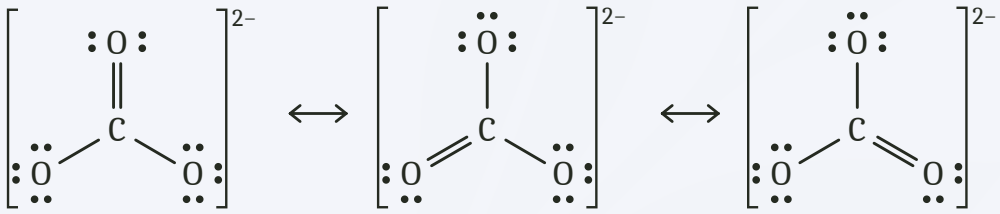
The ozone resonance hybrid.

5.3 Deducing resonance structures

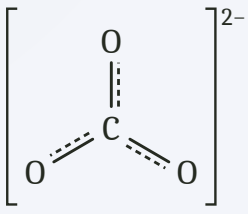
1. Draw one valid Lewis formula, counting valence electrons and adjusting for charge (Booklet 1).
2. Look for a double bond that could equally well be placed between the central atom and a **different, equivalent** outer atom.
3. Draw one structure for each possible position of the double bond. The **number of resonance structures equals the number of possible positions**. Move only electrons, never atoms, and keep every structure's brackets and charge.

Worked example — the carbonate ion, CO_3^{2-}

Valence electrons: 4 + (3 × 6) + 2 = **24**. Carbon needs one C=O double bond to complete its octet, and that double bond can be placed to any of the three oxygen atoms, so there are **three** resonance structures:

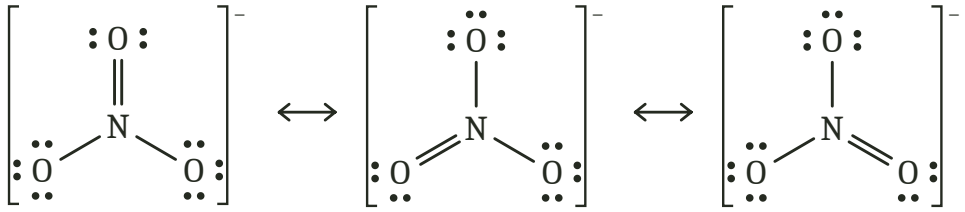


In the hybrid, the four bonding pairs and the extra delocalized pair are spread equally over all three C–O bonds. All three bonds are identical, with a length between those of C–O (143 pm) and C=O (122 pm). Each bond is in effect "one and a third" bonds, and each oxygen carries two-thirds of the 2− charge.

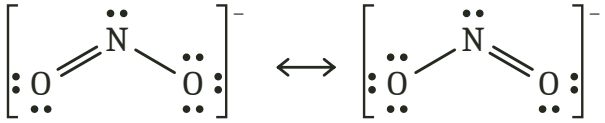


Other species with resonance

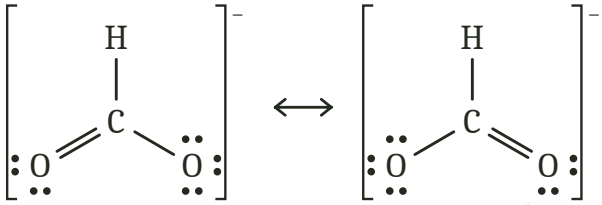
Nitrate ion, NO_3^- — 5 + (3 × 6) + 1 = 24 electrons; three resonance structures, like carbonate.



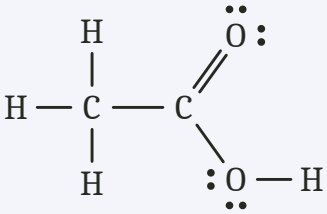
Nitrite ion, NO_2^- — 5 + (2 × 6) + 1 = 18 electrons; two resonance structures, with a lone pair on N.



Methanoate ion, $HCOO^-$ — 1 + 4 + (2 × 6) + 1 = 18 electrons; two resonance structures.

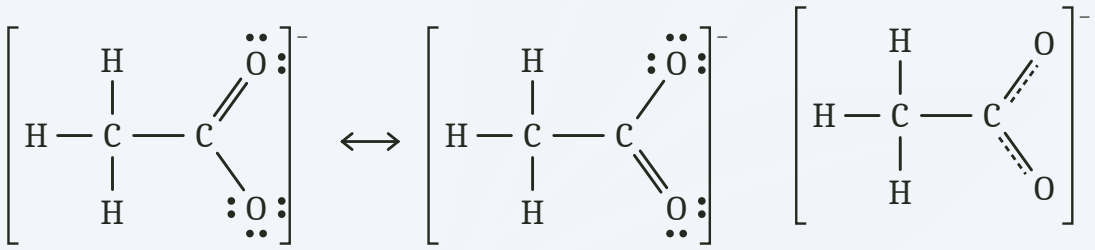


Worked example — compare the bonding in ethanoic acid, CH_3COOH, and the ethanoate ion, CH_3COO^-



Ethanoic acid has only one possible Lewis formula. The two oxygen atoms are not equivalent — one carries the H — so the double bond has only one possible position. The molecule has one C=O bond (122 pm) and one longer C–O bond (143 pm).

The ethanoate ion has two equivalent oxygen atoms, so the double bond has two possible positions and there are two resonance structures. The real ion is the hybrid: both C–O bonds are identical, intermediate between single and double, and the negative charge is shared equally by the two oxygens.



hybrid

Spreading the charge over two oxygen atoms stabilizes the ethanoate ion. This is part of the reason carboxylic acids are acidic while alcohols, whose anions cannot delocalize their charge, are not. Resonance therefore affects bond lengths, bond strengths and reactivity.

5.4 Oxygen, ozone and ultraviolet light

This is the Structure 1.3 linking question. A bond breaks when it absorbs a photon with at least the bond enthalpy (per molecule). The O=O bond in O_2 (498 kJ mol^{-1}) is stronger than the resonance-hybrid bond in O_3 (about 364 kJ mol^{-1}), so O_2 needs higher-energy, shorter-wavelength light to dissociate.

Worked example — the maximum wavelength that can break each bond

Energy per molecule, E = bond enthalpy ÷ N_A, and λ = hc / E.

O_2: E = 498 000 / 6.02×10^{23} = 8.27×10^{-19} J; λ = ($6.63 \times 10^{-34} \times 3.00 \times 10^{8}$) / 8.27×10^{-19} = 2.40×10^{-7} m = **240 nm**

O_3: E = 364 000 / 6.02×10^{23} = 6.05×10^{-19} J; λ = 3.29×10^{-7} m = **329 nm**

Both lie in the ultraviolet. Ozone absorbs the longer-wavelength UV (below about 330 nm) that oxygen cannot, which is why the ozone layer protects life from harmful UV-B radiation. The constants h, c and N_A are given in the data booklet (Section 2).

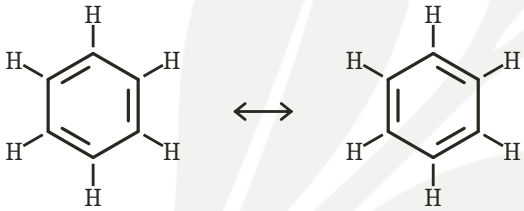
Examiner traps in 2.2.11

- Describing resonance as the molecule "flipping" between structures.
- Moving atoms between resonance structures. Only the positions of electrons (the double bond and lone pairs) change.
- Drawing a different number of electrons in each structure, or forgetting brackets and charge on an ion.
- Drawing resonance structures for a species such as ethanoic acid, CH_3COOH, where the outer atoms are not equivalent.
- Leaving the word "delocalized" out of an explanation. The guide requires it.

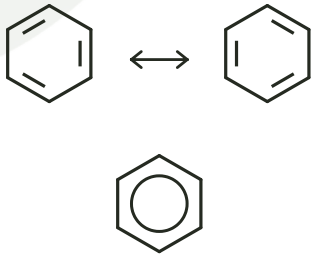
6 Structure 2.2.12 — Benzene HL

6.1 The Kekulé structure and its problems

Benzene, C_6H_6, has a carbon-to-hydrogen ratio of 1 : 1, which suggests a highly unsaturated molecule with several multiple bonds. In 1865 August Kekulé proposed a ring of six carbon atoms with **alternating single and double bonds** (cyclohexa-1,3,5-triene). Each carbon is bonded to one hydrogen and has three electron domains, so the molecule is planar with 120° bond angles. The double bonds can be placed in two ways, giving two resonance structures:



In skeletal form, where each corner is a carbon atom and the hydrogens are not shown, these become the two Kekulé hexagons. The true structure is the **resonance hybrid**: the six electrons of the three "double bonds" are delocalized evenly around the whole ring. This is drawn as a hexagon with a **circle** inside it.



benzene (hybrid)

Do not forget the circle

A hexagon **without** the circle is **cyclohexane**, C_6H_{12}, a completely different, saturated compound. When you draw benzene, always include the ring.

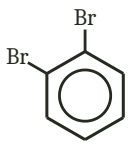


cyclohexane

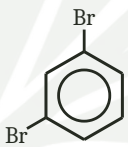
6.2 The evidence for delocalization

Kekulé's model explained benzene's formula and symmetry, but several observations do not fit alternating single and double bonds. Together, these form the physical and chemical evidence the guide asks you to discuss.

Evidence	Observation	Explanation
1. Bond lengths (physical)	X-ray diffraction shows all six C–C bonds are the **same length**, 139 pm, between C–C (154 pm) and C=C (134 pm). The ring is a regular hexagon.	Each bond has an equal share of the delocalized electrons, so all six bonds are identical and intermediate between single and double — in effect "one and a half" bonds. Kekulé's structure would have three short and three long bonds.
2. Bond enthalpy (physical)	The C–C bond enthalpy in benzene is about 507 kJ mol^{-1}, between C–C (346) and C=C (614).	The same averaging of bond character.
3. Enthalpy of hydrogenation (thermochemical)	$C_6H_6 + 3H_2 \rightarrow C_6H_{12}$ releases 208 kJ mol^{-1}, far less than the 3 × 120 = 360 kJ mol^{-1} expected for three C=C bonds.	Benzene is more stable than the Kekulé structure by about **152 kJ mol^{-1}**. This is the **resonance (delocalization) energy**: delocalization lowers the energy of the electrons.
4. Type of reaction (chemical)	Benzene does not readily undergo the addition reactions typical of alkenes (for example, it does not decolourize bromine water at room temperature). It undergoes **substitution** reactions instead.	Addition would destroy the delocalized ring and lose the resonance energy. Substitution replaces an H atom while keeping the stable ring intact.
5. Number of isomers (chemical)	Only **three** isomers of dibromobenzene, $C_6H_4Br_2$, exist.	With six identical bonds there are only three possible arrangements. Kekulé's structure would predict two different 1,2-isomers: one with the Br atoms across a single bond and one across a double bond.



1,2-dibromobenzene



1,3-dibromobenzene



1,4-dibromobenzene

6.3 The resonance energy and benzene's unreactivity

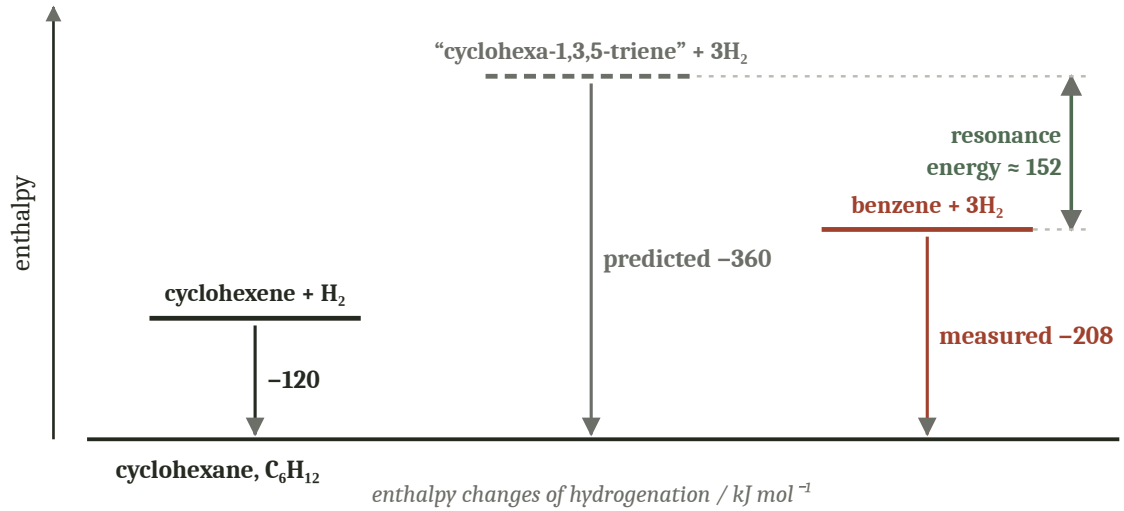


Figure 4 — Enthalpy changes of hydrogenation. Benzene is about 152 kJ mol^{-1} lower in energy than a ring of three isolated C=C bonds.

This answers the Reactivity 2.1 and 2.2 linking question. Hydrogenating cyclohexene releases 120 kJ mol^{-1} for one C=C bond, so three isolated C=C bonds should release about 360 kJ mol^{-1}. Benzene releases only 208 kJ mol^{-1}, because it starts about 152 kJ mol^{-1} lower in energy. Any reaction that destroys the delocalized ring must supply this resonance energy as well, which raises the energy barrier, so benzene is much less reactive towards addition than an alkene.

6.4 Why benzene undergoes electrophilic substitution

This answers the Reactivity 3.4 linking question. The ring of six delocalized electrons lies above and below the plane of the carbon atoms, forming a region of high electron density. That electron density attracts **electrophiles** (electron-pair acceptors). But if the electrophile were simply added, the delocalized ring would be lost. Instead, the electrophile replaces a hydrogen atom, so the product keeps the stable delocalized ring. Benzene therefore favours **electrophilic substitution** over addition. The mechanisms are covered in Reactivity 3.4.

HL Each carbon in benzene is sp^2 hybridized (2.2.16). Three sp^2 orbitals form σ bonds in the plane; the remaining p orbital on each carbon overlaps sideways with both neighbours, forming a continuous π system above and below the ring.

Nature of Science — a model revised by new evidence

Kekulé's model was a major advance and explained much of what was known in 1865. It had to be refined when new techniques produced data it could not explain: X-ray diffraction measured six equal bond lengths, and thermochemistry measured a hydrogenation enthalpy far smaller than predicted. Science progresses this way — a model is kept while it fits the evidence and modified, or abandoned, when it does not. The willingness to revise a well-established model in the light of new data is part of what makes scientific knowledge reliable.

Model answer: "Explain why all the carbon–carbon bonds in benzene are the same length." (2 marks)

The six electrons of the π bonds are delocalized evenly around the whole ring, rather than confined to three alternate double bonds **[1]**. Each carbon–carbon bond therefore has the same share of electrons, so all six bonds are identical and intermediate in length between a single and a double bond **[1]**.

Examiner traps in 2.2.12

- Drawing benzene as a plain hexagon (cyclohexane) or as a single Kekulé structure when the question asks for the true structure.
- Saying benzene has "three double bonds and three single bonds".
- Quoting the resonance energy without explaining where it comes from: the experimental hydrogenation enthalpy is less exothermic than predicted.
- Saying benzene is unreactive because its bonds are "strong double bonds". It is the stability of the **delocalized** ring that resists addition.

7 Bringing it all together SL + HL

One chain of reasoning for any covalent substance

Lewis formula (Booklet 1) → **shape** and **polarity** (Booklet 2) → **intermolecular forces**: London always; dipole–dipole if polar; hydrogen bonding if H is on N, O or F → **physical properties**: boiling point, volatility, solubility, R_F → compare only at similar molar mass.

HL Where a double bond can sit in more than one equivalent position, draw the **resonance structures**; the real species is a **hybrid** with delocalized electrons, equal bonds and extra stability.

Substance	M / g mol^{-1}	Strongest intermolecular force	Volatility	Soluble in water?	R_F on silica, hexane solvent
hexane, C_6H_{14}	86	London	high	no	high
propanone, CH_3COCH_3	58	dipole–dipole	high	yes (accepts H bonds from water)	intermediate
ethanoic acid, CH_3COOH	60	hydrogen bonding	low	yes	low

8 Key points to memorise

Everything below is either wording that examiners look for, or knowledge that is *not* in the data booklet. If you learn nothing else from this booklet, learn this section.

A Definitions — near-verbatim

- **London (dispersion) force:** attraction between an instantaneous dipole in one molecule and the dipole it induces in a neighbour. Present between all molecules; increases with the number of electrons.
- **Dipole–dipole force:** attraction between the permanent dipoles of polar molecules.
- **Dipole–induced dipole force:** attraction between a permanent dipole and the dipole it induces in a non-polar molecule.
- **Van der Waals forces:** the collective term for London, dipole–induced dipole and dipole–dipole forces.
- **Hydrogen bond:** attraction between an H atom covalently bonded to N, O or F and a lone pair on an N, O or F atom of a neighbouring molecule.
- **Volatility:** the tendency of a substance to vaporize.
- **R_F:** distance travelled by the component ÷ distance travelled by the solvent, both from the origin.
- HL **Resonance structures**, **resonance hybrid** and **delocalization** (section 5.2).

B Strength order and how to use it

- At comparable molar mass: **London < dipole–dipole < hydrogen bonding**. All are far weaker than covalent or ionic bonds.
- London forces increase with more electrons and with less branching (more surface contact).
- Melting and boiling a simple molecular substance overcomes intermolecular forces, not covalent bonds.
- Anomalously high boiling points: NH_3, H_2O, HF — hydrogen bonding. Ice is less dense than water: open, hydrogen-bonded tetrahedral lattice.

C Solubility and conductivity

- Like dissolves like: polar in polar (water), non-polar in non-polar (hexane).
- Solubility of alcohols in water falls as the hydrocarbon chain gets longer.
- Simple covalent substances do not conduct, unless they ionize in water (HCl).

D Chromatography

- Stationary phase: water in the cellulose of paper, or silica / alumina in TLC. Mobile phase: the solvent.
- Pencil origin above the solvent; lid on; mark the solvent front at once.
- R_F between 0 and 1, no units. High R_F: prefers the mobile phase. On silica with a non-polar solvent, polar components have low R_F.

E HL Resonance and benzene

- Number of resonance structures = number of possible positions for the double bond. Species to know: O_3 (2), CO_3^{2-} (3), NO_3^- (3), NO_2^- (2), $HCOO^-$ (2), CH_3COO^- (2), C_6H_6 (2).
- Hybrid: equal bonds, intermediate length and strength, charge shared, extra stability; no lone pairs drawn.
- Benzene: C–C 139 pm (C–C 154, C=C 134); resonance energy ≈ 152 kJ mol^{-1} (360 predicted − 208 measured); substitution not addition; three dibromobenzene isomers.

F Command terms — what each one is actually asking for

- **Deduce / Identify** (forces) — name *every* force present, starting with London.
- **Explain** (a boiling point difference) — name the forces in **both** substances, say which is stronger, and link stronger forces to more energy needed.
- **Calculate** (R_F) — show the division; no units.
- **Interpret** (R_F) — relate the value to the relative attraction of the component for the two phases.
- HL **Discuss** (benzene) — give several pieces of evidence, each with the observation and its explanation.

9 Check yourself

Cover the answers on the next pages. Marks are shown in brackets where a question is written in exam style. Questions marked HL are for Higher Level students.

1. Describe how London (dispersion) forces arise between molecules of bromine, Br_2. **[3]**
2. State all the intermolecular forces present between the molecules of: (a) $CO_2(s)$ (b) $NH_3(l)$ (c) $N_2(l)$ (d) $CH_3OCH_3(l)$ (e) $CHCl_3(l)$ (f) a mixture of $CH_3OCH_3(l)$ and $C_6H_{14}(l)$
3. Draw a diagram to show a hydrogen bond between two molecules of methanol, CH_3OH. Include the relevant lone pair and partial charges. **[2]**
4. Which substance has the highest boiling point?

   A $CH_3CH_2CH_3$ B CH_3OCH_3 C CH_3CH_2OH D CH_3CHO
5. For each pair, state which has the lower boiling point and explain why: (a) C_2H_6 and C_3H_8 (b) H_2O and H_2S (c) Cl_2 and Br_2 (d) HF and HCl (e) pentane and 2,2-dimethylpropane
6. Explain why the boiling points of the group 16 hydrides increase from H_2S to H_2Te, and why H_2O does not fit this trend. **[3]**
7. Explain why ice floats on liquid water. **[2]**
8. Which substance is the most soluble in water?

   A CH_3OH B CH_4 C C_2H_6 D $C_7H_{15}OH$
9. Explain why ethanol is soluble in water but hexane is not. **[3]**
10. Pure hydrogen chloride does not conduct electricity, but a solution of hydrogen chloride in water does. Explain. **[2]**
11. Five solids, P to T, are an ionic compound, a non-polar molecular solid, a polar molecular solid, a metal and a covalent network. Identify each.

Solid	Solubility in water	Conductivity of solid	Conductivity of aqueous solution	Melting point
P	insoluble	good	—	high
Q	soluble	none	good	high
R	insoluble	none	—	very high
S	soluble	none	none	low
T	insoluble	none	—	very low

12. Which two measurements are needed to calculate the R_F value of a spot?

    A height of origin; height of spot B distance from origin to solvent front; distance from origin to spot

    C length of paper; distance from origin to spot

    D distance from bottom of paper to solvent front; distance from bottom of paper to spot
13. In a paper chromatogram, the solvent front was 12.0 cm from the origin, and two spots were 9.0 cm and 6.0 cm from the origin. Calculate their R_F values. **[2]**
14. A mixture of two compounds, X and Y, was separated by TLC on silica using a non-polar solvent. X travelled 24 mm, Y travelled 72 mm and the solvent front 80 mm. (a) Calculate the R_F values of X and Y. (b) Outline why Y travelled further. **[4]**

15. State two practical precautions when setting up a paper chromatogram, and the reason for each. **[2]**

16. HL Draw the two resonance structures of the nitrite ion, NO_2^-, and the resonance hybrid. **[2]**

17. HL Put these species in order of increasing carbon–oxygen bond length, and explain your order: CO, CO_2, CO_3^{2-}, CH_3OH. **[3]**

18. HL Compare the nitrogen–oxygen bond lengths in the nitrate ion, NO_3^-, and in nitric acid, HNO_3 (structure HO–NO_2), with reference to resonance. **[3]**

19. HL Which of these is evidence for delocalization in benzene? I. all C–C bond lengths are equal II. benzene is reluctant to undergo addition reactions III. only three isomers of dibromobenzene exist

A I and II only B I and III only C II and III only D I, II and III

20. HL The enthalpy change of hydrogenation of cyclohexene is −120 kJ mol^{-1} and that of benzene is −208 kJ mol^{-1}. Use these values to estimate the resonance energy of benzene, and explain what it shows. **[3]**

10 Answers

1. The electrons in a Br_2 molecule are constantly moving, so at an instant the electron density may be uneven, producing an instantaneous dipole **[1]**. This induces a dipole in a neighbouring molecule by repelling or attracting its electrons **[1]**. The δ+ end of one dipole attracts the δ− end of the other **[1]**.

2. (a) London only (b) London, dipole–dipole and hydrogen bonding (c) London only (d) London and dipole–dipole (e) London and dipole–dipole (f) London, dipole–dipole (between CH_3OCH_3 molecules) and dipole–induced dipole (between CH_3OCH_3 and C_6H_{14}).

3. Diagram: O–H of one molecule with H δ+ and O δ−; a dashed line from that H to a lone pair on the O of a second methanol molecule **[1]**; O–H**-**O drawn in a straight line, with the lone pair shown **[1]**.

4. **C.** All four have similar molar mass; only ethanol forms hydrogen bonds.

5. (a) C_2H_6: fewer electrons, weaker London forces (b) H_2S: no hydrogen bonding, whereas water has hydrogen bonds (c) Cl_2: fewer electrons, weaker London forces (d) HCl: no hydrogen bonding, whereas HF has hydrogen bonds (e) 2,2-dimethylpropane: more branched and compact, less surface contact, weaker London forces.

6. From H_2S to H_2Te the number of electrons increases, so London forces get stronger **[1]**. H_2O has a much higher boiling point than predicted **[1]** because its molecules form hydrogen bonds, which are stronger than the forces between the other hydrides **[1]**.

7. In ice, each water molecule is hydrogen bonded to four others in a tetrahedral arrangement, forming an open lattice with a lot of space **[1]**. The molecules are further apart than in the liquid, so ice is less dense than water **[1]**.

8. **A.** Methanol is small and can hydrogen bond with water; heptan-1-ol has a long non-polar chain.

9. Ethanol has an –OH group, which can form hydrogen bonds with water molecules **[1]**. Hexane is non-polar and can only form London forces **[1]**, which cannot replace the hydrogen bonds between water molecules that must be broken for it to dissolve **[1]**.

10. Pure HCl consists of molecules, with no ions or delocalized electrons to carry charge **[1]**. In water, HCl ionizes to form $H^+(aq)$ (H_3O^+) and $Cl^-(aq)$ ions, which are free to move and carry charge **[1]**.

11. P metal; Q ionic compound; R covalent network; S polar molecular solid; T non-polar molecular solid.

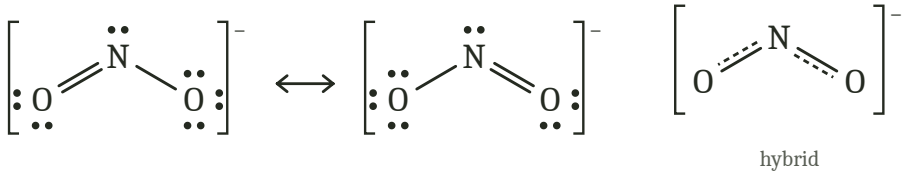
12. **B.**

13. 9.0 / 12.0 = **0.75 [1]**; 6.0 / 12.0 = **0.50 [1]**.

14. (a) X: 24 / 80 = 0.30 **[1]**; Y: 72 / 80 = 0.90 **[1]**. (b) Y is less polar than X, so it is less strongly attracted to the polar silica stationary phase **[1]** and more attracted to the non-polar mobile phase, so it moves further with the solvent **[1]**.

15. Any two: draw the origin in pencil, because ink would dissolve and run **[1]**; keep the origin above the solvent level, so the sample does not dissolve into the solvent **[1]**; use a lid, to keep the tank saturated with vapour so the solvent does not evaporate from the paper; mark the solvent front immediately, before the solvent evaporates.

16. Two structures, each with one N=O and one N–O, a lone pair on N, brackets and a 1− charge, linked by ↔ **[1]**. Hybrid: two identical N–O bonds each drawn as a solid line plus a dashed line, in brackets with the charge **[1]**.



17. $CO < CO_2 < CO_3^{2-} < CH_3OH$ **[1]**. CO has a triple bond and CO_2 double bonds **[1]**; in CO_3^{2-} each bond is a hybrid of about 1⅓ bonds, and CH_3OH has a C–O single bond. Fewer shared electrons between the atoms means a longer bond **[1]**.

18. In NO_3^- there are three resonance structures, so the electrons are delocalized and all three N–O bonds are identical, intermediate between single and double **[1]**. In HNO_3 the O bonded to H is not equivalent to the others, so its N–O bond is a single bond and is the longest **[1]**; the other two N–O bonds share one double bond between them by resonance, so they are equal and shorter **[1]**.

19. **D.** All three are consistent with delocalization and not with the Kekulé structure.

20. Predicted for three isolated C=C bonds: $3 \times (-120) = -360$ kJ mol^{-1} **[1]**. Resonance energy = 360 − 208 = 152 kJ mol^{-1} **[1]**. Benzene is 152 kJ mol^{-1} more stable than the Kekulé structure because its electrons are delocalized **[1]**.

Where this booklet leads

- HL **Structure 2.2 — Booklet 4**: expanded octets (2.2.13), formal charge (2.2.14), sigma and pi bonds (2.2.15) and hybridization (2.2.16).
- **Structure 1.5** — real and ideal gases.
- **Structure 3.2** — functional groups, homologous series and their boiling points; infrared spectroscopy.
- HL **Reactivity 3.4** — electrophilic substitution of benzene.

These notes are a free resource, shared for any IB student who finds them useful.

If a particular idea here still isn't clicking, specialist one-to-one support in Physics, Chemistry and Maths from a past A Level examiner is available at iBLaurel — call or message **+44 7404 710097**, or use the contact form at **www.iBLaurel.com**.

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