INTERMOLECULAR FORCES

- Exist between molecules which are covalently bonded
- They are weak forces of attraction between molecules
- They are much weaker than actual covalent bonds
- Physical properties of simple covalent molecules depend on the intermolecular forces present in the molecule

Covalent Compounds

Simple Molecular or simple molecular lattice

Giant Covalent
- giant covalent lattice
- graphite, diamond

Polar covalent
- H_2O, ethanol
- $CHCl_3$
- CH_3COOH
- C_2H_3Cl
- NH_3

Non-polar covalent
- Cl_2, F_2, I_2, O_2
- any diatomic HC's
- C_2H_4
- CO_2

Q. what holds the H_2O molecule together? (Intermolecular forces)

Three main types of intermolecular forces:
1. Permanent Dipole – Permanent Dipole
2. Instantaneous Dipole – Induced Dipole
3. Hydrogen bonds

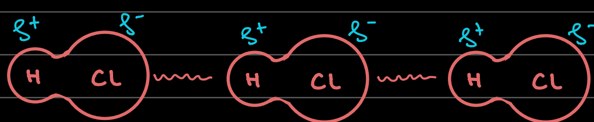
} Van der Waal's / London Dispersion forces

1. Permanent Dipole – Permanent Dipole forces (PD-PD)

- these occur in molecules that have a permanent dipole

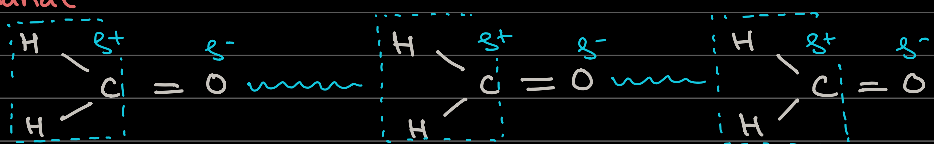
ie. $\overset{\delta+}{H} - \overset{\delta-}{Cl} \sim\sim \overset{\delta+}{H} - \overset{\delta-}{Cl} \sim\sim \overset{\delta+}{H} - \overset{\delta-}{Cl}$ (PD-PD, PD-PD)

$\sim\sim$ = intermolecular forces



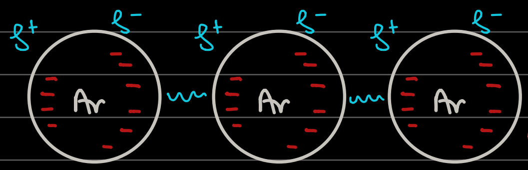
- Two molecules that have a permanent dipole will attract each other
- The molecules "flip" to give this arrangement

Methanal



- PD-PD exists in polar molecules (not in water → exception)
  ↳ dipoles do not cancel out

2. Instantaneous Dipole – Induced Dipole (van der waals)



↳ weakest forces of attraction

Fig 1.1

- Present in non polar molecules (also present in all covalent molecules, but let's ignore that for now)

Notes:

Instantaneous – Induced
- Electrons within a molecule are mobile and are constantly moving.
- This constant movement causes a fluctuating dipole

- At any one instant – a non-polar molecule, ie. Ar, can possess an instantaneous dipole (due to there being at that moment more electrons on one side of the molecule than the other).
↳ Refer to Fig 1.1

- This instantaneous dipole will induce a corresponding dipole in the neighboring molecule / atom.
- This results in an attraction between the molecules caused Instantaneous Dipole – Induced Dipole force of attraction (aka. Van der waal's forces)

- These ID-ID forces are present in all covalent molecules in addition to any other force of attraction.

ie. HCl → PD-PD (most dominant)
but it will also have ID-ID (much weaker)

* But explain the physical properties. ie. boiling point, in terms of the most dominant intermolecular force (ie. in this case, PD-PD)

- The ease of distortion of the electron cloud and therefore the strength of the ID-ID forces depends on:

① the number of electrons within a molecule / The number of polarisable electrons or Mr of the molecule → Higher the Mr, more polarisable electrons
② Molecular Shape

↳ Explaining these factors:

1. Number of electrons
- The greater the number of polarisable electrons → larger Mr / bigger molecule higher the boiling point. ← the stronger the ID-ID forces

* When some compound / molecule is boiled (liquid → gas), only the intermolecular

forces are broken and **NOT** the covalent bonds

Example: In group 18 noble gases → the higher the Mr → stronger ID-ID → higher boiling point

Gas	B.P (°C)
He	-269
Ne	-249
Ar	-186
Kr	-152
Xe	-108
Rn	-61.8

Q. Explain the trend.

A. As the Mr increases, there are more polarisable electrons, which makes the ID-ID forces stronger and more energy is required to break them and hence, the boiling points increase.

Example: Hydrides of Group VI

Hydride	B.P (°C)
CH_4	-162
SiH_4	-112
GeH_4	-90
SnH_4	-52

Q. Explain the trend.

Same explanation applies here

2. Shape of the molecule

↳ Essentially looking at if a molecule is straight vs. branched

*Isomers: compounds with the same molecular formula but a different structural formula

Butane C_4H_{10}

1. $CH_3CH_2CH_2CH_3$ → Straight chain isomer → B.P = 36°C

2. $CH_3 - C(CH_3)(H) - CH_3$ → branched isomer → B.P = 9°C

Straight Chain

- A straight chain molecule can exist closer together / have a larger surface area

$CH_3CH_2CH_2CH_3$

$CH_3CH_2CH_2CH_3$

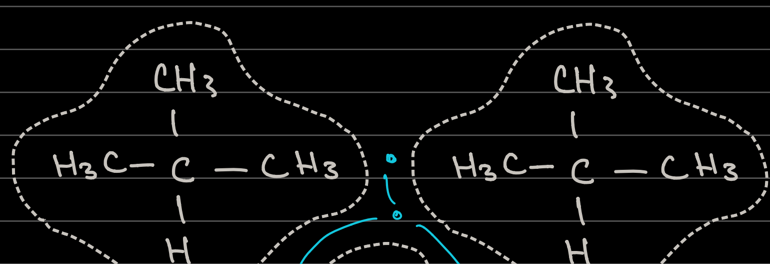
$CH_3CH_2CH_2CH_3$

more points of contact

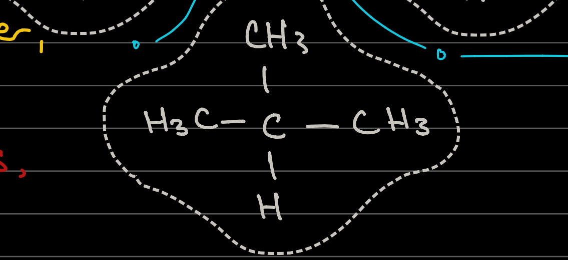
- There are more points of contact
- This makes ID-ID forces stronger
  ↳ leads to it having a higher boiling point

Branched

- More spherical
- fewer points of contact
- they cannot pack closer together



↳ This makes ID-ID forces weaker, hence a lower boiling point



fewer points of contact

* The greater the number of branches, the lower the boiling point

3. Hydrogen Bonds : strongest intermolecular force

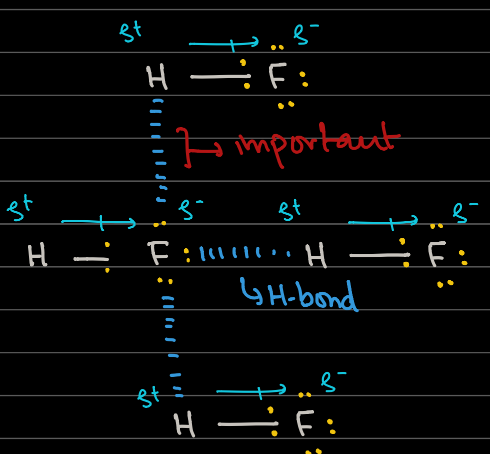
$\underbrace{\text{Covalent bonds}}_{\text{Chemical bond}} > \underbrace{\text{H bonds} > \text{PD-PD} > \text{ID-ID}}_{\text{Intermolecular bonds}}$

- Hydrogen bonds are forces of attraction between an electron deficient H bonded to a very electronegative atom and the lone pairs of electrons on a neighboring, very electronegative atom

- H-bonds are formed when the H is <u>directly</u> bonded to the electronegative atom (ie. F, O, N)

$^{\delta+}H - F^{\delta-}$
$^{\delta+}H - O^{\delta-}$ → electron-rich, very electronegative atom.
$^{\delta+}H - N^{\delta-}$

↳ These are the electron deficient H's, due to the electrons being pulled away by the highly electronegative atoms.



<u>Rules (of showing H-bonds)</u>
1. Show lone pairs
2. Show δ^+ and δ^-
3. Show H-bond as a dotted bonds from the LP to the electron deficient H

<u>Other tips to remember</u>:
1. Draw the diagram large and clear
2. Draw no more than 2 H-bonds, to maintain clarity

Q. Why do H's bonded to F, O, N only form H-bonds?

radius decreases →

1:46:20

P1	H							He	
P2	Li	Be	B	C	N	O	F	Ne	↓ radius increases
P3	Na	Mg	Al	Si	P	S	Cl	Ar	

	EN
✓ F	= 4.0
✓ O	= 3.5
✓ N	= 3.0
× Cl	= 3.0

N, O, F have the smallest radii, therefore, their charge density is the highest

↳ "highly" electronegative when bonded to H

↓

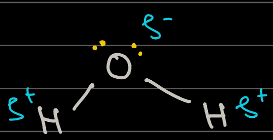
This causes H to become "highly" electron-deficient, which allows H-bonds to happen.

Q. Why not chlorine?

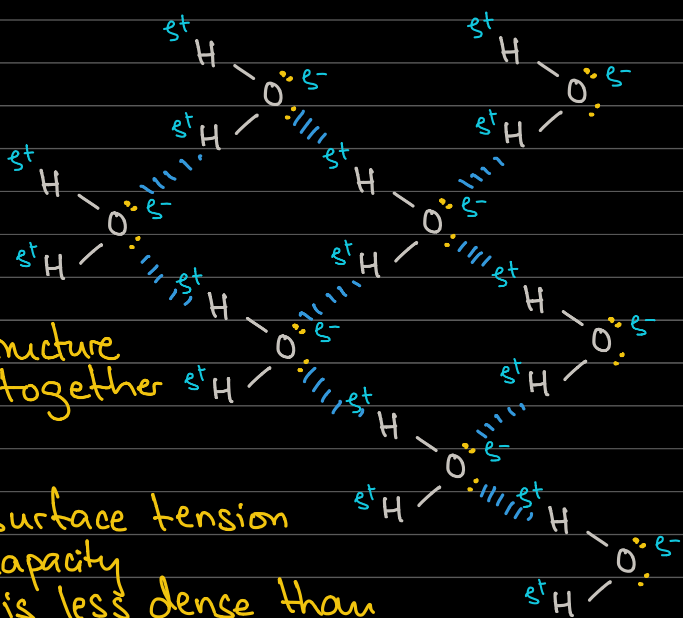
A. Cl has a larger covalent radius, so the H bonded to the Cl is not as electron deficient as the H bonded to F, O, or N.

As a result, H-Cl has only PD-PD forces of attraction and not H-bonds

Hydrogen bonding (in water) → important



- water has very extensive hydrogen bonds

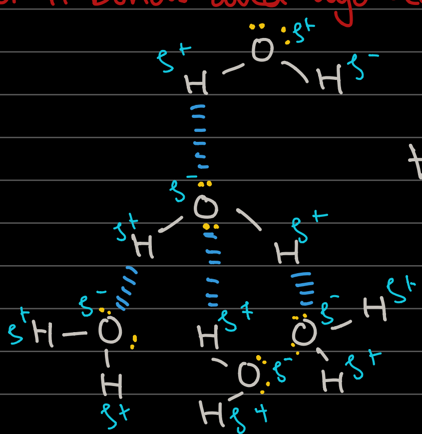


- Water has a huge lattice structure of water molecules bonded together by hydrogen bonds

this explains this →

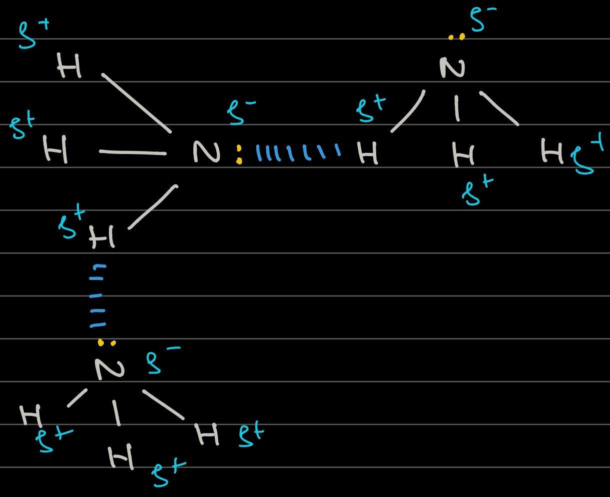
- High BP of 100°C, surface tension
- higher specific heat capacity
- Density of ice, ice is less dense than liquid water

▷: 2:06:00 for H-bonds and hydrocarbon molecules



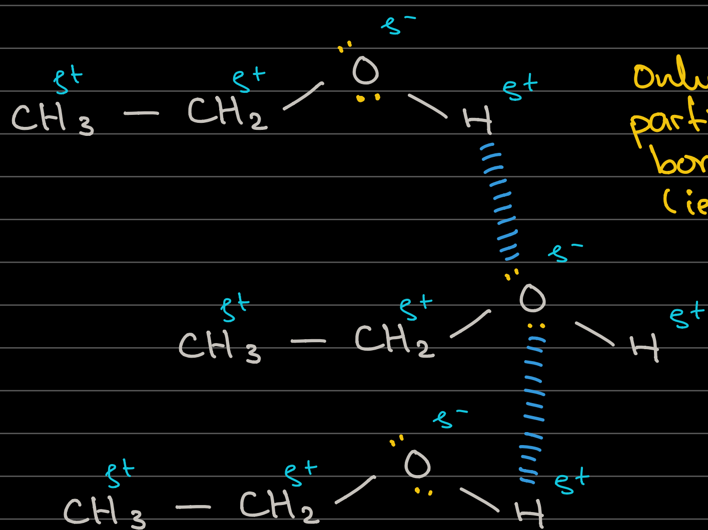
tetrahedral.

NH_3 (Ammonia) :



→ hydrogen bonding in NH_3 gas.

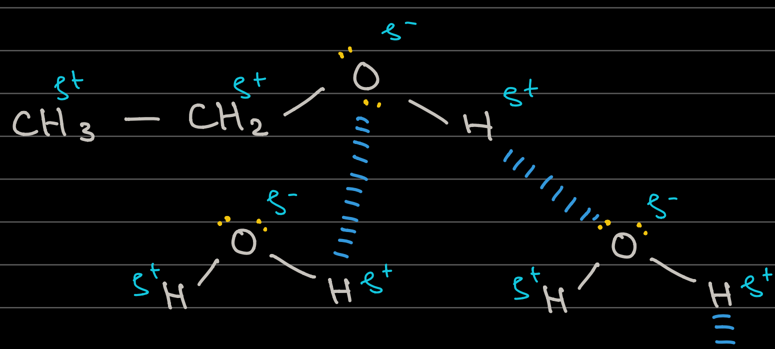
CH_3CH_2OH (ethanol) : molecules are held together by H-bonding

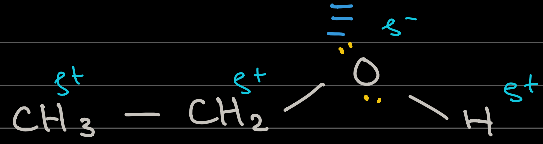


Only the O – H hydrogen takes part in H bonding, as it is directly bonded to an electronegative atom (ie. oxygen)
↪ this is the electron deficient H

- B.P of ethanol = 78°C , not very high, as the molecules are held together by strong H- bonds

Q. Why is ethanol soluble in water ?
A. They are both polar and can form hydrogen bonds with each other.



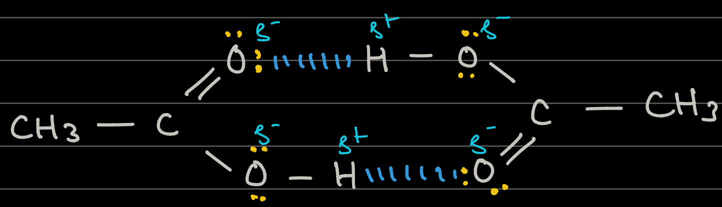


Two different types of H-bonds:
1. water's $O^{\delta-}$ to ethanol's $H^{\delta+}$
2. Ethanol's $O^{\delta-}$ to water's $H^{\delta+}$

2:40:00

CH_3COOH (ethanoic acid) pure (ie. not in water)
- Pure organic acids can be crystallized
- exits as white crystals

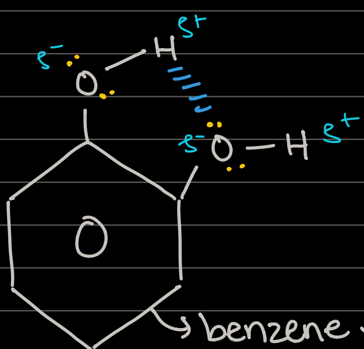
- Ethanoic acid can form "dimers" by H-bonding



→ ethanoic acid dimer
aka. "glacial acetic acid"
because it fine ice-like "needles" or crystals

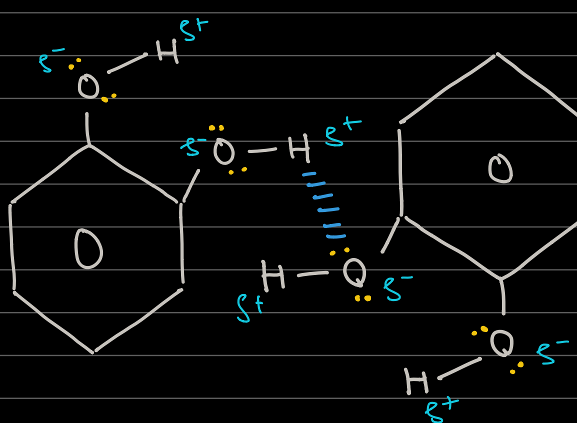
*Intramolecular: within the same molecule

Intramolecular H-bonding is possible if there are polar groups in close proximity



benzene ring

→ Intramolecular H-bonding



→ Intermolecular H-bonding for the same molecule

Another example of H-bonds :