LEAVING CERTIFICATE

CHEMISTRY NOTES

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2017

These notes are compiled based on the State Examinations Commission (SEC) syllabus with reference to teacher guidelines. All areas covered correlate with past exam papers (including 1990s) and some sample pre/mock papers. The "higher order questions" are designed to anticipate future questions which bring together multiple parts of the course to test understanding and reasoning. Marking schemes are also linked in to relevant parts with special attention on common student mistakes and how to achieve full marks in each question.

Unit 2 – Chemical Bonding

2.1 Chemical Compounds

A compound: is a pure substance consisting of two or more different chemical elements that can only be separated into similar substances by chemical reactions

Group	Ions Formed
1	+1
2	+2
3	+3
4	–
5	–1
6	–2
7	–3
8	–

Simple Chemical Formulae

Examples:

1) Sodium Chloride: Na^{+} + Cl^{-} ⟶ NaCl

2) Aluminium Oxide: Al^{+3} + O^{-2} ⟶ Al_2O_3

Complex Ions

Examples:

Ion	Formula
Carbonate	CO_3^{2-}
Hydrogen Carbonate	HCO_3
Sulfate	SO_4^{2-}
Sulfite	SO_3^{2-}
Nitrate	NO_3^{-}
Nitrite	NO_2^{-}
Phosphate	PO_4^{3-}
Hydroxyl	OH^{-}
Ammonium	NH_4

1) Calcium Carbonate: Ca^{+2} + CO_3^{2-} ⟶ $CaCO_3$

2) Iron(II) Nitrate: Fe^{2+} + NO_3^{-} ⟶ $Fe(NO_3)_2$

3) Iron (III) Nitrate: Fe^{3+} + NO_3^{-} ⟶ $Fe(NO_3)_3$

Iron, a transitional metal, has varying valencies – iron (II) and iron (III)

Noble Gases

Reference unit 1 for chemical and physical properties

- Group 0/8
- All elements have a full outer shell (energy levels)
- As a result of **full energy levels** they are **very stable** and **unreactive (inearth)**

Practical Uses of Noble Gases

Helium and argon do not form any compounds due to their unreactivity

1. Helium is a much safer alternative to hydrogen e.g. balloons, blimps (both low densities, helium not flammable)
2. Electric light bulbs contain argon

Transition Metals

Elements 21 ⟶ 29 (Sc ⟶ Cu)

Element	Valency
Cu	1,2
Cr	2,3,6
Fe	2,3,6
Mg	2,3,4,6,7

- Variable valency e.g. Iron(II) and iron (III)
- Produced coloured ions
- Often used as catalysts
- Form complex ions e.g. haemoglobin
- 2 electrons in outer shell ($4s^2$) and incomplete 3^{rd} shell

Note:

Zn: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$

Forms Zn^{2+} ions → Valency = 2

Bonding and Valency

Valency: the number of bonds an atom of the element forms when it reacts

- **Note:** Valency is not the same as oxidation number (no charges)

COMPOUND	FORMULA	VALENCE	OXIDATION STATE
Hydrogen chloride	HCl	H=1 Cl=1	H= +1 Cl= −1
Chlorine	Cl_2	Cl=1 Cl=1	Cl= +1 Cl= −1

Octet Rule:

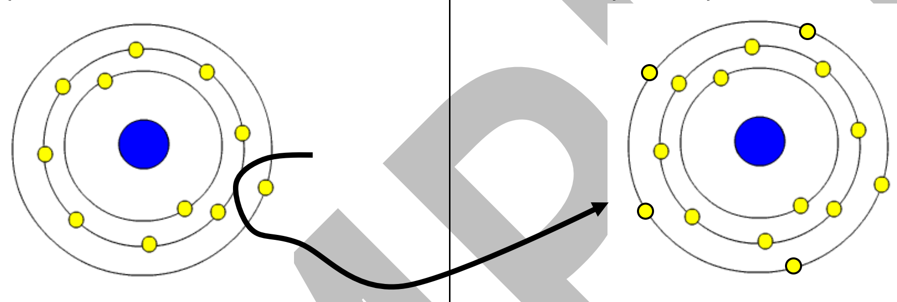
Elements combine to form compounds and their atoms react in such a way as to achieve 8 electrons in their outer shell

Sodium

Na: $1s^2\,2s^2\,2p^6\,3s^1$

Chlorine

Cl: $1s^2\,2s^2\,2p^6\,3s^2\,3p^5$



- Sodium has 1 electron in its outer shell/energy level
- Chlorine has 7 electrons in its outer shell/energy level
- When sodium and chlorine react: sodium loses its electron and chlorine gains it

New electronic configurations:

Na^{+1}: $[1s^2\,2s^2\,2p^6]^+$ Sodium now has 8 electrons in outer shell – stable

Cl^{-1} : $[1s^2\,2s^2\,2p^6\,3s^2\,3p^6]^-$ Chlorine now has 8 electrons in outer shell – stable

Limitations:

1. Does not work on the first four elements – impossible to gain/loses electrons
2. Hydrogen and lithium gain/lose to get to He but they are unlikely gain electrons to reach 8 electrons in the outer shell
3. Be and B have few electrons in their outer shell so it is unlikely that they will reach 8 electrons in the outer shell
4. "d-block" elements excluded

2.2 Ionic Bonding

Intramolecular Bonding

Bonding **within** a molecule that holds that atoms together

<u>Types of Intramolecular Bonding:</u>

1. Ionic
2. Pure Covalent
3. Polar Covalent

Ionic Bonding

Electrostatic force of attraction between oppositely charged ions

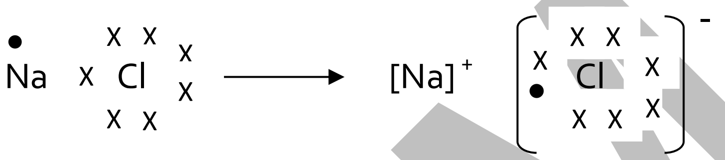
- Ions are formed due to transfer of electrons
- A bond involving transfer (loss <u>and</u> gain) of electrons – one molecule loses and electron; one molecule gains an electron
- Minute size of ions

Cations: are positive ions formed when an atom looses electrons (lose of negative charge – more positive)

Anions: are negative ions formed when an atom gains electrons (gain of positive charge – more negative)

Representation of Ionic Bonds Using Dot and Cross Models

Sodium Chloride (NaCl) *Note: only outer shell electrons are represented!*

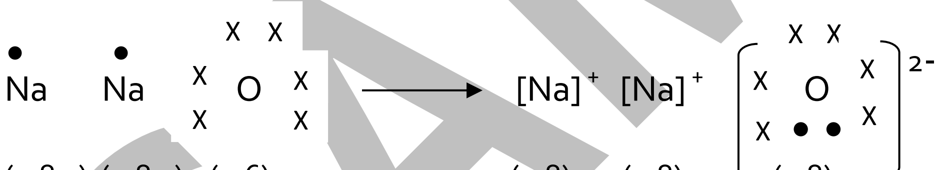


(2,8,1) (2,8,7) (2,8) (2,8,8)

Na: $1s^2 2s^2 2p^6 3s^1$ → Na^{+1}: $[1s^2 2s^2 2p^6]^{+1}$

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$ → Cl^{-1}: $[1s^2 2s^2 2p^6 3s^2 3p^6]^{-1}$

Sodium Oxide (Na_2O)



(2,8,1) (2,8,1) (2,6) (2,8) (2,8) (2,8)

Na: $1s^2 2s^2 2p^6 3s^1$
O: $1s^2 2s^2 2p^4$
→
Na^{+1}: $[1s^2 2s^2 2p^6]^+$
O^{2-}: $[1s^2 2s^2 2p^6]^{2-}$

Na^{+1}: $[1s^2 2s^2 2p^6]^+$
Cl^{-1}: $[1s^2 2s^2 2p^6 3s^2 3p^6]^-$
O^{2-}: $[1s^2 2s^2 2p^6]^{2-}$
All 3 now have full outer "shells" which confers extra stability

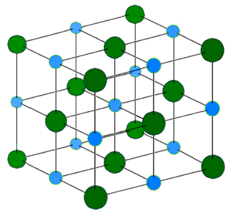
Characteristics of Ionic Substances

1. Do not conduct electricity in **solid form** as the ions cannot move
2. Can conduct electricity in **liquid/molten form** as the ions can move freely
   - Most ionic substances dissolve in water
3. Solids at room temperature
4. High melting and high boiling point
   - Strong forces between ions (intermolecular) mean a lot of energy is required to break them up

Ionic Materials in Everyday Life

1. Salt tablets to replace salt lost by sweat
2. Bleach – Sodium Hypochlorite

Sodium Chloride Crystal Structure Between Molecules



- Electrostatic force of attraction between sodium and chloride ions is not just one to one but occurs in all directions around the ion
- Therefore, an ionic compound such as NaCl consists not just of a pair of ions but a network of ions held together in a regular and repeating pattern

2.3 Covalent Bonding

Bond formed due to **sharing of electrons**

- **No ions formed**
- Each pair of electrons shared is a single covalent bond

Molecule: 2 or more atoms chemically combined

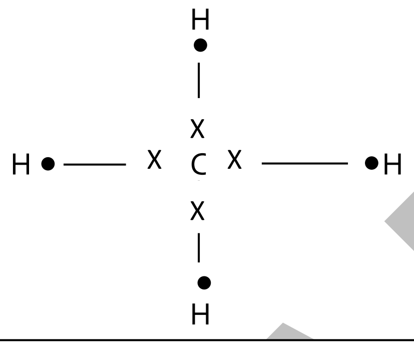
- Minute size

Representation of Covalent Bonds Using Dot and Cross Models

Methane (CH_4)

CH_4 = for every C we need 4 H

C: $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1$ H: $1s^1$



Typical Exam Question:

Description of the bond:

- All bonds are single covalent bonds
- Carbon has 4 bond pairs and no lone pairs
- Hydrogen has 1 bond pair and no lone pairs

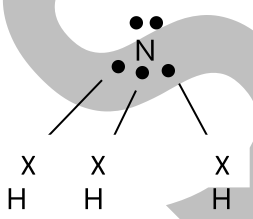
Rule for lone pairs:

Group 5 – 1 lone pair
Group 6 – 2 lone pairs

Bond pairs: shared electrons involved in bonding (outer shell electrons)

Lone pairs: shared electrons not involved in bonding

Ammonia (NH_3)



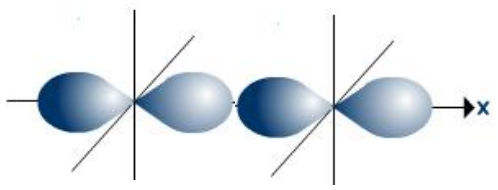
Description of the bond:

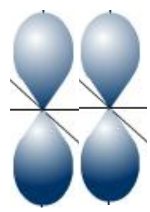
- All bonds are single covalent bonds
- Nitrogen has 3 bond pairs and 1 lone pair
- Each Hydrogen has 1 bond pair and no lone pairs

Types of Covalent Bonds

Sigma (σ) Bonds: covalent, head on overlap of atomic orbitals

Pi (π) Bonds: covalent, side on overlap of P atomic orbitals





Nitrogen

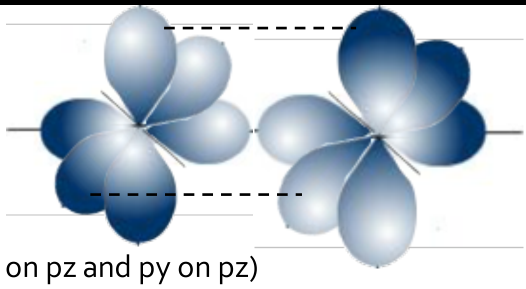
$1s^2\ 2s^2\ \underline{2px^1\ 2py^1\ 2pz^1}$

Each N shares 3 electrons

1 sigma bond (head on head)

2 pi bonds (side on side – dashed lines, pz on pz and py on pz)

1 lone pair (not shown – $2s^2$; $1s^2$ 2 electrons not represented as only the outer shell is represented)



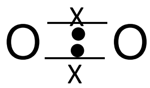
$N \equiv N$

Non-Polar (Pure) Covalent Bonding

Bond formed due to equal sharing of electrons
- No partial charges formed

- The majority of non-polar covalent bonds occur in diatomic elements
- Other examples include Methane (CH_4)

Example: O_2



1 bond pairs and 2 lone pairs are on each oxygen (not shown)

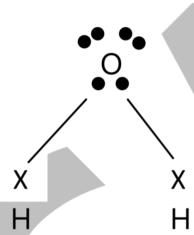
Since electronegativity of the 2 oxygen are equal no partial charges are formed and the electrons are shared equally

Polar Covalent Bonding

Bond formed due to unequal sharing of bonding electrons
- ✓ Partial positive (δ^+) and partial negative (δ^-) charges formed

H_2O

H: $1s^1$ O: $1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$

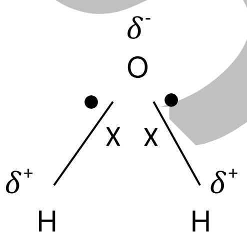


H^+, O^{2-} i.e. for every O we need 2 H

But this representation is not the reality.

Oxygen has a higher electronegativity value that hydrogen leaving the following representation: see later

[See notes 2.4 for electronegativity]



Characteristics of Covalent Substances

1. Gases/Liquids at room temperature
2. Low boiling and melting points (due to weak intermolecular forces)
3. Do not conduct electricity
4. Insoluble in water

Everyday Materials – *Syllabus Examples*

Polar:	**Non-Polar:**
1. Petrol	1. Water
2. Cooking oil	2. Glucose

2.4 Electronegativity

The relative attraction of an atom for a shared pair of electrons in a covalent bond

Trends

Reference: Log Book pg 81

Across a period: Increase

- Decreasing atomic radius
- Increase in effective nuclear charge (more electrons)

Down a group: Decrease

- Increasing atomic radius
- Increase in nuclear charge but this is offset by screening effect

Prediction of Bond Type Using Electronegativity Differences

Value	Bond Type
< 1.7	Polar Covalent
> 1.7	Ionic
≈ 0 (0-0.4)	Non-Polar Covalent

Example: H_2O

Electronegativity values: H = 2.20 and O = 3.44

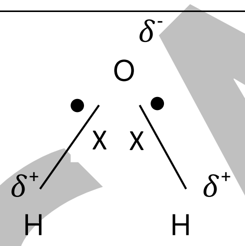
O – H (oxygen minus hydrogen)

3.44 – 2.20

1.24

∴ **H_2O is Polar Covalent**

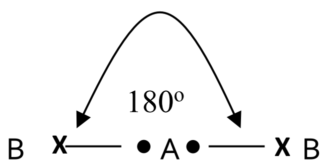
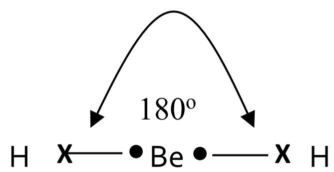
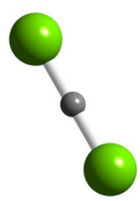
Note:
Only take one atom of each element and subtract them from each other

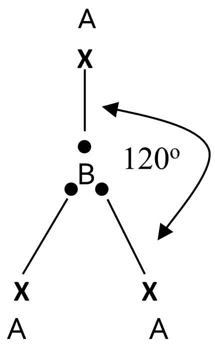
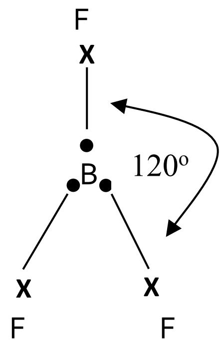
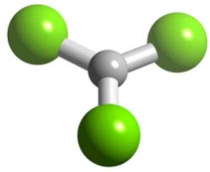
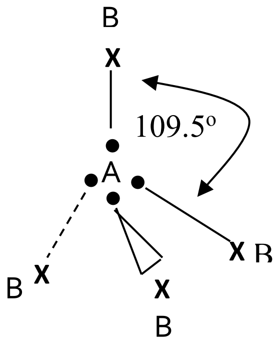
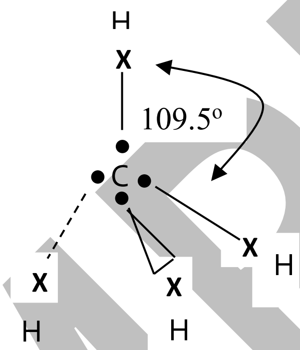
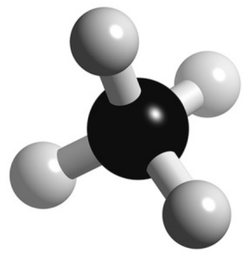


- ✓ H_2O is polar covalent
- ✓ Unequal sharing of electrons
- ✓ Pair of electrons closer to the oxygen, matching electronegativity values

2.5 Shapes of Molecules and Intermolecular Forces

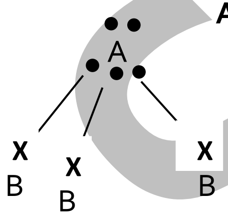
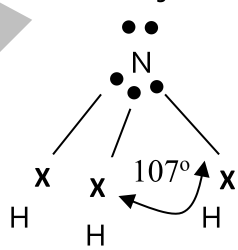
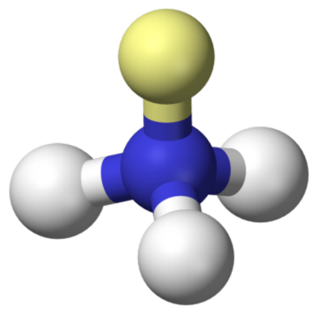
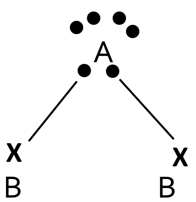
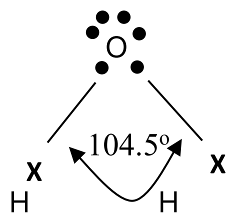
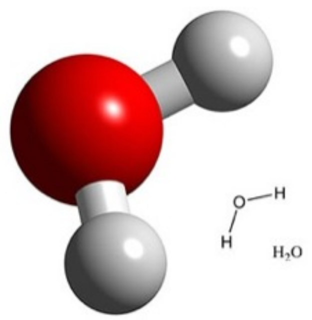
Simple Molecules

GENERAL SHAPE	EXAMPLES	3-D IMAGE
LINEAR 2 Bond Pairs **AB_2** 	**BeH_2** 	

TRIGONAL PLANAR 3 Bond Pairs **AB_3** 	**BF_3** 	
TETRAHEDRAL 4 Bond Pairs **AB_4** 	**CH_4** 	

Complex Molecules

Electron pair repulsion theory: the electrons pairs in the valence (outer) shell of the central atom repel each other and end up as far apart as geometrically possible

GENERAL SHAPE	EXAMPLES	3-D
PYRAMIDAL 3 Bond Pairs + 1 Lone Pair **AB_3**  A has to be in group 5 - 1 lone pair	**NH_3** 	
V-SHAPED 2 Bond Pairs + 2 Lone Pairs **AB_2**  A has to be in group 6- 2 lone pairs	**H_2O** 	

Using Electron Pair Repulsion Theory to Explain Shapes of Molecules

Typical Exam Question – example: Describe the shape of ammonia

1. **N has 3 bonding** and **1 non-bonding** (lone) pair (Each H has 1 bond pair)
2. The bond arrangement causes the shape of the molecule to be **pyramidal**

Electron Pair Repulsion Theory

Bond Angles

1. AB_3 with no lone pairs has a bond angle of 120°
   AB_3 with 1 lone pair has a bond angle of 107°
2. AB_2 with no lone pairs has bond angle of 180°
   AB_2 with 2 lone pairs has bond angle of 104.5°

Theory

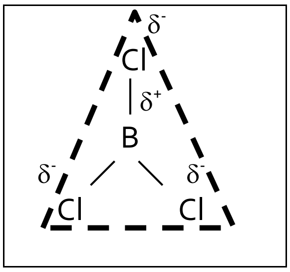
- Lone pairs have greater repulsion of each other – i.e. lone pair in contact with lone pair will produce the greatest repulsion, followed by lone pair in contact with bond pair and finally the weakest, bond pair in contact with bond pair
- L.P:L.P > L.P:B.P > B.P:B.P (Repulsion) **
- Lone pair wants to get as far apart as geometrically possible from each other pushing bonds closer together **

** **Both points are required to get full marks**

Relationship between Symmetry and Polarity in a Molecule

B – Cl

- Electronegativity difference = 1.12
- Therefore one may assume it is polar
- However, BCl_3 is actually non-polar
- This is due to unequal sharing of electrons between B and Cl (i.e. polarity) cancels due to symmetry of molecule
- Centres of positive and negative charges coincide
- BCl_3 has a trigonal planar shape which has symmetry



Intermolecular Forces

Attractive/Repulsive attractive forces **between** molecules

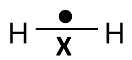
Types of Intermolecular Forces:

1. Van der Waal's forces
2. Dipole-dipole
3. Hydrogen Bonding

Van der Waal's Forces

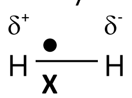
Very weak intermolecular forces

Up until now H_2 was represented as:

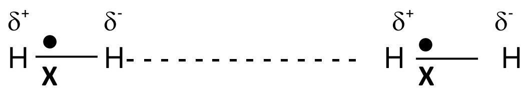


giving rise to a pure covalent intramolecular bond

But, in reality the 2 electrons in H_2 are moving from side to side creating temporary charges



creating temporary polarity, known as **a temporary dipole** within the molecule



The greater number of electrons in a molecule, the greater number of possible temporary dipoles, and therefore the greater intermolecular attraction

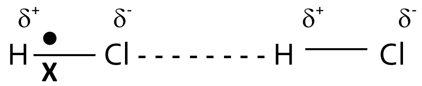
This means that Van der Waal's forces increase with an increasing size of molecule – i.e. bigger molecule has more electrons

These weak intermolecular forces increase the boiling point with the more temporary dipoles

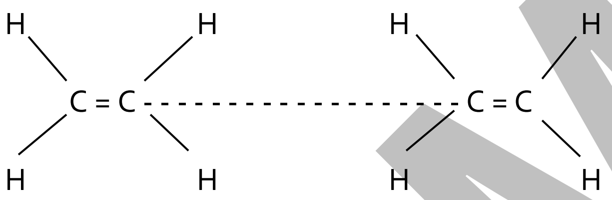
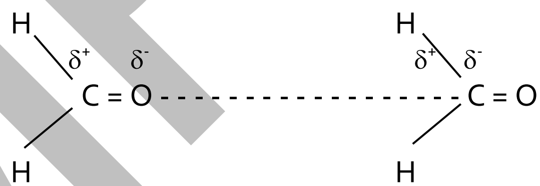
E.g. Oxygen (16 electrons) has a much higher boiling point than Hydrogen (1 electron) – *syllabus*

Dipole-dipole

- Intermolecular forces between polar molecules
- Differ from Van der Waal's forces **by permanent dipoles due to the polarity** of the molecule



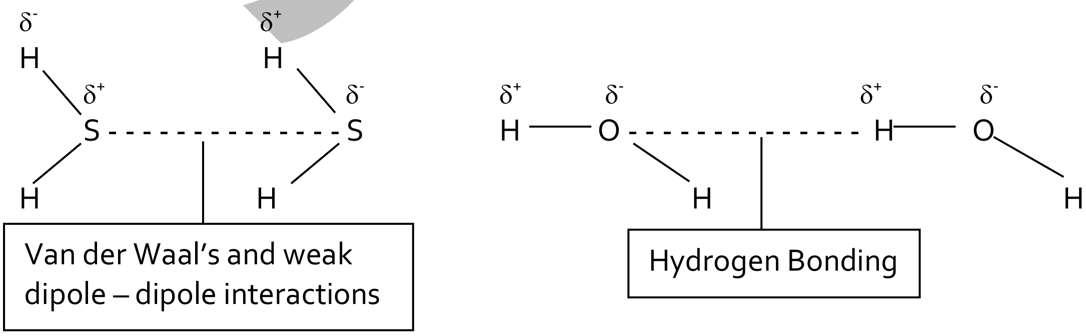
- Due to permanent dipole the boiling point of molecules with dipole-dipole nteractions are much higher than molecules with Van der Waal
- ***Syllabus:*** **Ethene C_2H_4 (Mr=28) should have similar boiling point to Methanal HCHO (Mr=30), however Methanal has a much higher boiling point due to stronger intermolecular bonding**

Ethene contains Van der Waal's forces only:	Methanal contains dipole-dipole forces only:
	

Hydrogen Bonding

Intermolecular attraction involving a slightly positive hydrogen atom bonded to a **small highly** electronegative element such as F, O or N

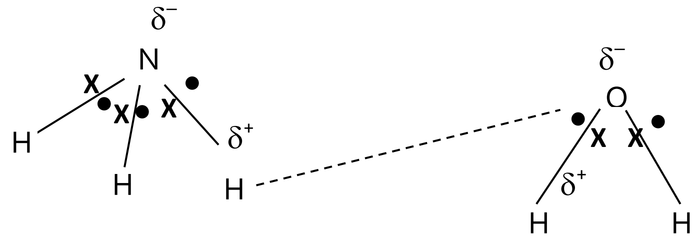
- Hydrogen bonding is the strongest form of intermolecular bonding
- This is because the **molecules are highly polar**
- E.g. in water molecules, O – H is highly polar (large electronegativity value)
- **H_2S should have a higher boiling point to water due to greater relative molecular mass. But since the H – S bond is less polar than the O – H bond in water it has a much lower boiling point than water – *syllabus example***



H_2S has an electronegativity difference of 0.38 which means it is between polar and non-polar

Application of Knowledge: Dissolving Properties

Ammonia and Water:



Process

- The slightly negative O in water bonds with the H in ammonia
- The slightly negative N in ammonia bonds with H in water
- Breaking of hydrogen bonds in water
- Forming of hydrogen bonds between ammonia and water

Note: Even though both compounds contain hydrogen bonding, water has a much higher boiling point because of larger electronegativity difference in the OH bond than the NH bond of ammonia

2012: state how bonding in PH_3 differs from NH_3, H_2O, HCl

- PH_3 = non-polar
- NH_3, H_2O, HCl = polar

Reason for this difference in bonding?

- Tiny electronegative different in PH_3
- Large electronegative difference in the others
- As the bonding type gets stronger:
  - Increase in boiling point
  - Increase in melting point

2.6 Oxidation Numbers

Oxidation Numbers: The charge that an atom appears to have when the electrons are distributed according to certain rules

Rules for Oxidation Numbers

1. Free elements have oxidation number of 0
   - e.g. N_2
2. The sum of the oxidation numbers is 0
   - H_2O

     $2(+1) - 2 = 0$
3. The oxidation number of a simple ion is equal to the charge of the ion (see 2.1 for ions)
   - $Cl^- = -1$
4. The sum of the oxidation numbers of all atoms in a complex ion is equal to the charge on the ion
   - $(NO_3^-) = -1$
5. **Hydrogen has an oxidation number of +1 in its compounds, except in metallic hydrides (hydrogen + metal) where it is – 1**

6. **Oxygen has an oxidation number of – 2 in its compounds, except in hydrogen peroxide where it is – 1, and when bonded to fluorine, where it is + 2**

Can check which element is more electronegative in log tables. The more electronegative element is assigned oxidation numbers first.

Naming of Transition Metal Compounds

➢ **Note:** If 2 elements only in a compounds, we use the prefix –ide; –ide ending means no oxygen

Sample questions:

What is the systematic name of $FeCl_3$?

Iron (II) chloride

What is the formula of iron (III) sulfate-9-water?

Fe^{3+}, SO_4^{2-}, $9H_2O$ = $Fe_2(SO_4^{2-})_3.9H_2O$

Oxidation and Reduction in Terms of Oxidation Numbers

Oxidation: is an increase in oxidation number – loss of electrons

Reduction: is a decrease in oxidation number – gain of electrons

Oxidising Agent: is a substance that cause reduction

Reducing Agent: is a substance that cause oxidation

Oxidation
Is
Loss
Reduction
Is
Gain

➢ Bleaches as examples of oxidising agents (e.g. NaOCl) or reducing agents (e.g. SO_2)

Example:

What is (a) oxidised (b) reduced (c) the oxidising agent (d) the reducing agent in the following redox reaction?

$2MnO_4^-$	+	$16H^+$	+	$5C_2O_4^{2-}$	→	$2Mn^{2+}$	+	$10CO_2$	+	$8H_2o$
+7 **-2**		**+1**		+3 **-2**		**+2**		+4 **-2**		+1 **-2**
↑				↑				↑		↑
+7 *-8*		*+1*		+6 *-8*		+2		+4 -4		+2 -2

1. Assign oxidation numbers according to previous rules – **in bold**
2. Now account for number of elements in compounds – *in italic* e.g. O_4 – oxygen has oxidation number of -2 but since there is 4 of them it becomes -8 (note: ignore number of moles e.g. $4O_4$ is still -8)
3. Work out oxidation numbers for the rest of the compound noting the overall charge – underlined e.g. CO_2, O: -2 x 2 = -4 therefore C must have +4 charge as the overall charge of CO_2 is neutral; MnO_4^-, O: -2 x 4 = -8 therefore M must have **+7 and not + 8** because the **overall charge of MnO_4^- is -1**
4. Assign individual molecule oxidation numbers; e.g. H_2O - O: -2 x 1 = -2, H: +2 (overall charge is neutral) but since there is 2H (H_2) molecules **each individual H has +1** – represented by arrows

Mn: +7 → +2 **Gain of 5 electrons (5 negative charges)** – REDUCED therefore OXIDISING AGENT

C: +3 → + 4 Loss of 1 electron **(1 negative charge)** – OXIDISED therefore REDUING AGENT

Balancing Redox Equations Using Oxidation Numbers

We use the "ratio" method to solve equations that cannot be balanced by inspection

Example:

MnO_4^-	+	CH_3OH	+	H^+	$\longrightarrow$	Mn^{2+}	+	HCHO	+	H_2O
+7 **-2**		-2 **+1 -2 +1**		**+1**		**+2**		**+1** 0 **+1 -2**		**+1 -2**
↑		↑ ↑						↑		↑
+7 -8		-2 +3 -2 +1		+1		+2		+1 0 +1 -2		+2 -2

1. Assign oxidation numbers according to previous rules – **in bold**
2. Now account for number of elements in compounds – *in italic* e.g. O_4 – oxygen has oxidation number of -2 but since there is 4 of them it becomes -8 (note: ignore number of moles e.g. $4O_4$ is still -8)
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4. Assign individual molecule oxidation numbers; e.g. H_2O - O: -2 x 1 = -2, H: +2 (overall charge is neutral) but since there is 2H (H_2) molecules **each individual H has +1** – represented by arrows
   - CH_3OH: H: +1 x 1 = +1, H_3: +1 x 3 = +3, O: -2 x 1 = -2 – charge so far is +4 – 2 = +2 → therefore C must be -2 to make overall charge 0

Mn: +7 → +2 **Gain of 5 electrons (5 negative charges)** – REDUCED therefore OXIDISING AGENT

C: -2 → 0 Loss of 2 electron **(2 negative charges)** – OXIDISED therefore REDUING AGENT ↖

5. Assign a ratio Mn : C - i.e. for every 2 Mn there is 5 C (reverse numbers to get ratio)
   2 : 5
6. Fill in equation according to ratio:

$$\mathbf{2}MnO_4^- + \mathbf{5}CH_3OH + H^+ \longrightarrow \mathbf{2}Mn^{2+} + \mathbf{5}HCHO + H_2O$$

7. Balance equation – count each element on each side of the equation:

Mn:	2	= 2	2	= 2	Balanced
O:	8 (2x4) + 5	= 13	5 (5 x 1) + 1	= 6	
C:	5	= 5	5	= 5	Balanced
H:	21 (3+1 x 5) + 1	= 21	12 (5 x 2)+ 1x2	= 12	

$$2MnO_4^- + 5CH_3OH + H^+ \longrightarrow 2Mn^{2+} + 5HCHO + \mathbf{8}H_2O$$

$8H_2O$ – O now balanced 13 on each side

This changes H - left side remains as 21, **right side now is 26**

Finally, balance H (always leave to last!) – **$6H^+$**

$$2MnO_4^- + 5CH_3OH + \mathbf{6}H^+ \longrightarrow 2Mn^{2+} + 5HCHO + 8H_2O$$

Now count each element again to ensure all balance)