VCE Chemistry Unit 1 Revision Notes + Practice Questions

SECTION A THEMES (Multiple Choice)

1. Atomic Structure and Isotopes

Revision Dot Points

- Atomic number = number of protons
- Mass number = protons + neutrons
- Isotopes have same number of protons but different neutrons
- Neutral atoms have equal protons and electrons
- Electrons occupy shells, subshells and orbitals
- Maximum electrons:
    - s = 2
    - p = 6
    - d = 10
    - f = 14
- Ions form by gain/loss of electrons

Practice Questions

Q1 How many neutrons are in $^{35}_{17}Cl$?

A. 17
B. 18
C. 35
D. 52

Neutrons = mass number − atomic number = 35 − 17 = 18.

Q2 Which pair are isotopes?

A. Na and Mg
B. ^{12}C and ^{14}C
C. Cl and Cl^-
D. O and O^{2-}

Isotopes => same number of protons different number of neutrons. Eg same atomic number different mass number

Q3 Which electron configuration belongs to Mg^{2+}?

A. $1s^2 2s^2 2p^6 3s^2$
B. $1s^2 2s^2 2p^6$
C. $1s^2 2s^2 2p^6 3s^1$
D. $1s^2 2s^2 2p^5$

Mg^{2+} loses 2 electrons leaving $1s^2$, $2s^2$, $2p^6$

2. Periodic Table Trends

Revision Dot Points

- Atomic radius decreases across a period
- Atomic radius increases down a group
- Ionisation energy increases across a period
- Metals lose electrons easily
- Non-metals gain electrons easily
- Electronegativity increases across a period

Practice Questions

Q1 Which element has the highest electronegativity?

A. Na
B. Mg
C. Cl
D. Ar

Chlorine has highest electronegativity

Q2 Which element has the largest atomic radius?

A. Li
B. Na
C. K
D. Rb

Rb has the largest atomic, as radius increases down group.

Q3 Which element is most metallic?

A. F
B. Cl
C. Na
D. O

Na

3. Ionic and Covalent Bonding

Revision Dot Points

- Ionic bonding = attraction between oppositely charged ions
- Covalent bonding = sharing electrons
- Metals + non-metals = ionic
- Non-metal + non-metal = covalent
- Ionic compounds:
    - high melting points
    - conduct electricity when molten or aqueous
    - brittle
- Polar bonds form due to electronegativity differences

Practice Questions

Q1 Which compound is ionic?

A. CO_2
B. NH_3
C. NaCl
D. H_2O

NaCl

Q2 Which bond is most polar?

A. H–H
B. Cl–Cl
C. O–H
D. C–C

O-H

Q3 Which property belongs to ionic compounds?

A. Low melting point
B. Conducts electricity as solid
C. Brittle
D. Insoluble in water

Brittle.
Not all ionic compounds are insoluble in water

4. Lewis Structures and Molecular Shape

Revision Dot Points

- Valence electrons determine bonding
- Single bond = 2 shared electrons
- Lone pairs affect shape
- Common shapes:
    - linear
    - bent (V-shaped)
    - pyramidal
    - tetrahedral
- Use VSEPR theory (simplified meaning = electron cloud, such as bonding and non-bonding pairs, repulsion in 3D space)

Practice Questions

Q1 What is the shape of NH_3?

A. Linear
B. Bent
C. Pyramidal
D. Tetrahedral

Pyramidal

Q2 How many lone pairs are on oxygen in H_2O?

A. 0
B. 1
C. 2
D. 3

2

Q3 What is the shape of CO_2?

A. Bent
B. Linear
C. Pyramidal
D. Tetrahedral

Linear

5. Intermolecular Forces and Boiling Point

Revision Dot Points

- London dispersion forces occur in all molecules
- Dipole–dipole forces occur in polar molecules
- Hydrogen bonding occurs when H bonded to N, O or F
- Stronger intermolecular forces → higher boiling point
- Larger molecules usually have stronger dispersion forces

Practice Questions

Q1 Which compound shows hydrogen bonding?

A. CH_4
B. HCl
C. NH_3
D. CO_2

NH_3

Q2 Which has the highest boiling point?

A. Methane
B. Ethanol
C. Ethanoic acid
D. Propane

Ethanoic acid

Q3 Why does water have a high boiling point?

A. Ionic bonding
B. Hydrogen bonding
C. Covalent bonding
D. Metallic bonding

Hydrogen bonding

6. Stoichiometry and Empirical Formula

Revision Dot Points

- Mole = amount of particles (6.02×10^{23})
- Avogadro constant = 6.02×10^{23}
- Empirical formula = simplest whole number ratio
- Molecular formula = actual number of atoms
- Steps for empirical formula:
    1. Assume 100 g
    2. Convert g to mol by dividing by molar mass
    3. Divide by smallest number of mol
    4. Multiply to whole numbers if needed

Practice Questions

Q1 How many moles are in 22 g of CO_2?

A. 0.25
B. 0.50
C. 1.0
D. 2.0

0.5

Q2 What is the empirical formula of C_2H_4?

A. CH
B. CH_2
C. C_2H_4
D. C_4H_8

CH

Q3 How many particles are in 0.1 mol?

A. 6.02×10^{22}
B. 3.01×10^{23}
C. 6.02×10^{23}
D. 1.20×10^{24}

6.01×10^{22}

7. Organic Chemistry and Naming

Revision Dot Points

- Alkanes: single bonds
- Alkenes: double bonds
- Alcohols contain –OH
- Carboxylic acids contain –COOH
- Prefixes:
    - meth = 1 carbon
    - eth = 2
    - prop = 3
    - but = 4
    - pent = 5
- Longest chain chosen first
- Priority functional group gets lowest numbered carbon
- Substituent groups are placed in alphabetical order

Practice Questions

Q1 What is the formula of pentane?

A. C_5H_{10}
B. C_5H_{12}
C. C_5H_8
D. C_5H_{14}

C_5H_{12}

Q2 What functional group is in ethanol?

A. Carboxyl
B. Amino
C. Hydroxyl
D. Carbonyl

Hydroxyl (OH)

Q3 What is the name of $CH_3CH(OH)COOH$?

A. 1-carboxypropan-2-ol
B. 2-hydroxypropanoic acid
C. 2-olpanoic acid
D. Butanoic acid

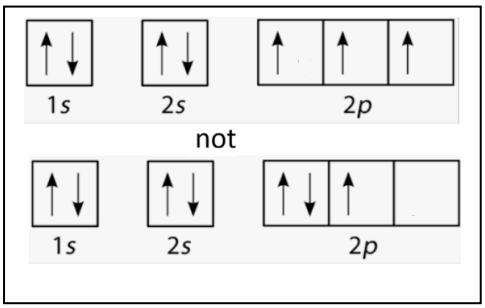
2-hydroxypropanoic acid

SECTION B THEMES (Short Answer)

1. Electron Configurations

Revision Dot Points

- Lowest energy levels fill first
- Orbitals Regions of space around the nucleus where electrons are likely to occupy) singly filled first , as shown on the right.



- Each orbital holds 0, 1 or 2 electrons (Pauli exclusion principle)
- Transition metals use d orbitals

Practice Questions

Q1 Write the electron subshell configuration for :

- sulfur._____ *$1s^2\,2s^2\,2p^6\,3s^2\,3p^4$*_____
- Copper _____*$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^1$*_____
  half filled subshell is more stable then partially filled. That is why 3d fills first with 4s half filled. More stable than $3d^9\,4s^2$. A full 3d is more stable than a full 4s

Q2 How many electrons fit in the 3d subshell?

_____*10*_____

Q3 Explain the exceptions to the dot point "Lowest energy levels fill first" and give an example.

Electrons generally fill the lowest energy orbitals first.

For example, within the same energy level, the s subshell fills before the p subshell because s orbitals are lower in energy.

Example:
$1s^2\,2s^2\,2p^2$ (Carbon)

In carbon, the 2s orbital fills completely before electrons enter the 2p orbitals.

Energy order:
1s < 2s < 2p < 3s < 3p

However, some atoms are exceptions to the normal filling order because half-filled and fully-filled subshells are more stable.

Example of an exception:
Copper = [Ar] $3d^{10}$ $4s^{1}$

Instead of $3d^{9}4s^{2}$, one electron moves from the 4s orbital into the 3d orbital so the 3d subshell becomes completely full and more stable.

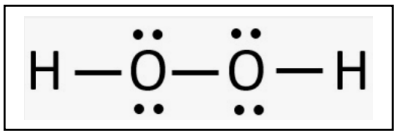
2. Lewis Structures and Shape

Revision Dot Points

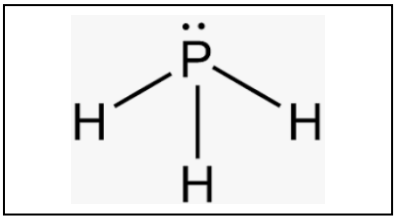
- Count valence electrons
- Central atom usually least electronegative
- Add lone pairs last
- Shape determined by bonding lone pairs

Practice Questions

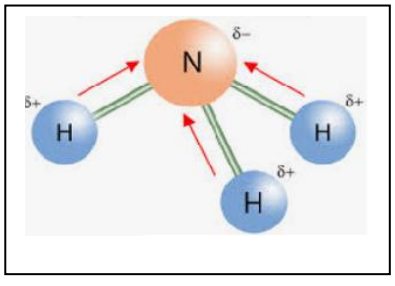
Q1 Draw the Lewis structure for H_2O_2.



Q2 Describe the shape of PH_3. Draw its structure



Q3 Explain why NH_3 is polar, draw the molecular structure.



NH_3 is a polar molecule where the dipole moments do not cancel out. There is a distinct δ+ and δ- end.

3. Percentage Composition and Mole Calculations

Revision Dot Points

- Percentage composition:

$$\% = \frac{\text{mass of element}}{\text{total molar mass}} \times 100$$

- Use molar masses correctly
- Convert between moles, mass and particles

Practice Questions

Q1 Find the percentage of oxygen in Al_2O_3.

Formula mass (Al_2O_3) = 102.0
Mass of oxygen = 48.0
=> % of O = 48.0 / 102 = 47.1%

Q2 Calculate the percentage composition of $CaCO_3$.

Step 1. Molar mass of $CaCO_3$

Total molar mass = 40.08 + 12.01 + 48.00 = 100.09 g/mol

Step 2. Percentage composition

Calcium (Ca):

$$\frac{40.08}{100.09} \times 100 = 40.04\%$$

Carbon (C):

$$\frac{12.01}{100.09} \times 100 = 12.00\%$$

Oxygen (O):

$$\frac{48.00}{100.09} \times 100 = 47.96\%$$

Ca = 40.04% : C = 12.00% : O = 47.96%

Q3 How many atoms are in 2 mol of H_2SO_4?

2 X 7 X 6.02 X 10^{23}

4. Ionic Compounds and Reactivity

Revision Dot Points

- Group 1 metals very reactive
- Reactivity increases down Group 1
- Metals lose valence electrons
- Ionic compounds conduct only when molten or in solution

Practice Questions

Q1 Why is potassium more reactive than lithium?

1. Atomic size increases down the group

Lithium atoms are smaller.

Potassium atoms are much larger (more electron shells).

2. Valence electrons are further from the nucleus

Both lithium and potassium have one valence electron.

In potassium, that electron is farther away and less strongly attracted to the nucleus.

3. Weaker nuclear attraction due to greater distance of valence electron from the nucleus hence potassium losses the electron with greater ease. More reactive.

Q2 State two properties of ionic compounds and explain your answer.

Any of the below.
- brittle – due to directional electrostatic forces between cations and anions in the crystal lattice.
- Conducts in the molten or aqueous state – due to mobile ions that are present liquid and aqueous states but not in the solid state.

- Relatively high meitng points -due to strong electrostatic forms present in the ionic crystal lattice.

Q3 Using the table of valencies found in your data book predict the formula of:

- magnesium fluoride. _____*MgF_2*_____

- ` potassium phosphate _____*K_3PO_4*_____

- Aluminium dichromate _____*$Al_2(Cr_2O_7)_3$*_____

5. Chemical Equations and Ionic Equations

Revision Dot Points

- Equations must balance
- Include states
- Spectator ions removed in net ionic equations
- Precipitates are insoluble solids (use the data book to gauge the solubility of compounds)

Practice Questions

Q1 Balance:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

Q2 Write the net ionic equation for silver nitrate reacting with sodium chloride.

_____*$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$* _____

Q3 What is a precipitate? Use the reaction between calcium nitrate solution and sodium phosphate solution, provide a balanced equation to support your answer.

Precipitates are insoluble solids formed from the mixing of two solutions. Calcium phosphate is a precipitate because it forms a solid when calcium nitrate solution is mixed with sodium phosphate solution.

$3Ca(NO_3)_2(aq) + 2Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + 6NaNO_3(aq)$

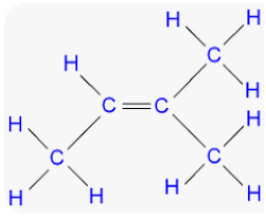
6. Organic Structures and Isomers

Revision Dot Points

- Structural isomers have same molecular formula (different connectivity)
- Draw structural formulae of organic compounds
- Draw condensed structural formulas carefully

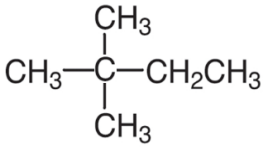
Practice Questions

Q1 Draw pent-2-ene.



$CH_3ChCH(CH_3)_2$

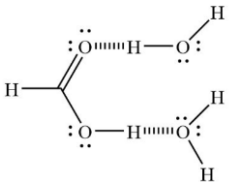
Q2 Name a structural isomer of hexane. Draw the structural formula of this isomer.

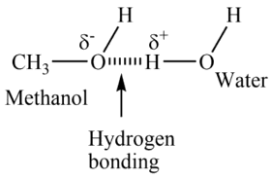


2,2-dimethylbutane

$C(CH_3)_3CH_2CH_3$

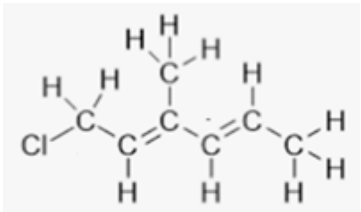
Q3 Explain why methanoic acid is more soluble than methanol. Draw a diagram to support your answer.





Methanoic acid can form two hydrogen bonds per molecule with water where as methanol can form only one has methanoic acid interacts more extensively with water.

Q4 Give the name and semistructural formula of the molecule shown on the right.



Name *1-chloro-3-methylhexa-2,4-diene* _

Semistructural formula

$CH_2ClCHC(CH_3)CHCHCH_3$

7. Empirical Formula Experiment

Revision Dot Points

A student determines the empirical formula of a metal oxide using a crucible experiment.

Results:

Item	Mass (g)
Crucible + lid	28.412
Crucible + lid + metal (before heating)	28.840
Crucible + lid + product (after heating)	28.921

The metal (X) has a molar mass of 56.0 g/mol and reacts with oxygen in air to form an oxide.

1. Calculate the empirical formula of the metal oxide. Show all working clearly.

Step 1: Calculate mass of metal before heating

Mass of metal = Mass of crucible + lid + metal − Mass of crucible + lid
= 28.840 g − 28.412 g = 0.428 g
Step 2: Calculate mass of metal oxide (product) after heating
Mass of metal oxide = Mass of crucible + lid + product − Mass of crucible + lid
= 28.921 g − 28.412 g = 0.509 g
Step 3: Calculate mass of oxygen combined
Mass of oxygen = Mass of metal oxide − Mass of metal
= 0.509 g − 0.428 g
= 0.081 g
Step 4: Calculate moles of metal and oxygen

- *Moles of metal (X) = Mass of metal / Molar mass of metal*
  = 0.428 g / 56.0 g/mol
  = 0.00764 mol
- *Moles of oxygen (O) = Mass of oxygen / Molar mass of oxygen*
  = 0.081 g / 16.0 g/mol
  = 0.00506 mol

Step 5: Determine mole ratio (metal : oxygen)

Divide both by the smaller number of moles (0.00506):

- *Metal: 0.00764 / 0.00506 ≈ 1.51*
- *Oxygen: 0.00506 / 0.00506 = 1*

Step 6: Convert ratio to whole numbers

Multiply both by 2 to remove decimal:

- *Metal: 1.51 × 2 = 3.02 (~3)*
- *Oxygen: 1 × 2 = 2*

Step 7: Write empirical formula

The mole ratio of metal to oxygen is approximately 3 : 2.

Empirical formula = X_3O_2

8. Chromatography and Solubility

Revision Dot Points

- More soluble in solvent => travels further
- Polar substances dissolve in polar solvents
- Rf value:

$$\text{Rf} = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

$$\text{Rf} = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

- Amino acids differ due to polarity

Practice Questions

Q1 What does a larger Rf value indicate?

A larger Rf value means the substance is more soluble in the solvent and travels further up the chromatography paper. This indicates it interacts less strongly with the paper and more strongly with the solvent.

Q2 Why do different amino acids separate?

Different amino acids separate because they have different polarities and solubilities. Polar amino acids dissolve better in polar solvents and travel differently

compared to non-polar amino acids. Their varying interactions with the stationary phase (paper) and mobile phase (solvent) cause them to move different distances.

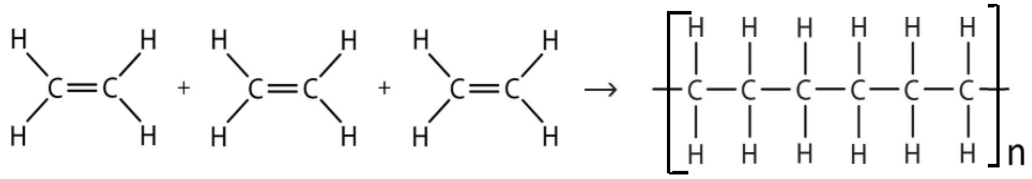
Q3 Would a polar amino acid travel further in water than hexane?

A polar amino acid would travel further in water because water is a polar solvent and it dissolves polar substances well. In hexane, which is non-polar, the polar amino acid would be less soluble and travel a shorter distance. Dispersion forces involved in hexane are not able to interact with H-bonding present in water.

9. Polymerisation

Revision Dot Points

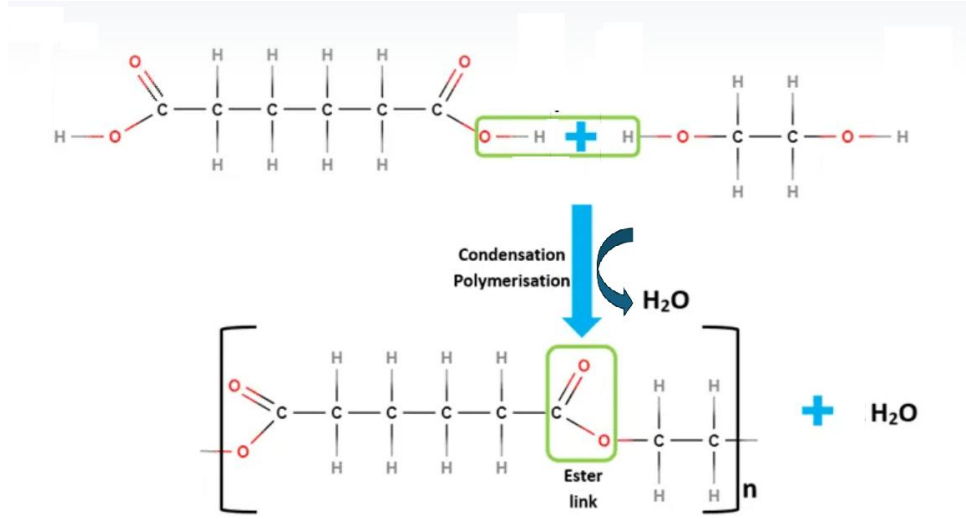
- Monomers are relatively small molecules that covalently bond together to form large molecules called polymers.
- Addition polymerisation occurs between unsaturated molecules containing carbon–carbon double bonds (C=C).
- In ad***dition polymerisation***, monomers join by breaking the C=C double bond, and only the carbon atoms originally part of the double bond form the backbone of the polymer chain.



Formation of polyethene from ethene monomers

- ***Condensation polymerisation*** occurs between molecules that have at least two reactive functional groups.

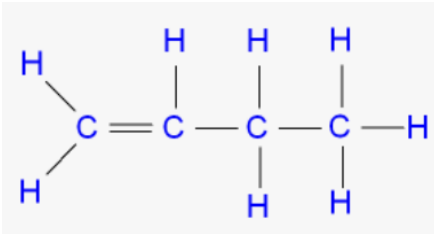
- The functional groups from each monomer react to form a covalent bond between monomers, releasing a small molecule such as water (H_2O) or hydrogen chloride (HCl) as a byproduct.



Summary of condensation polymerisation

Practice questions

Q1 . Consider the monomer shown on the right.

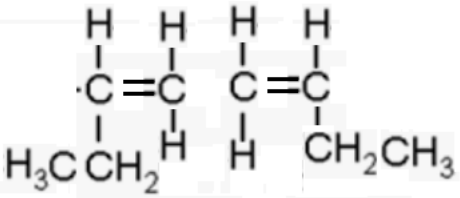


- Give the systematic name of the hydrocarbon

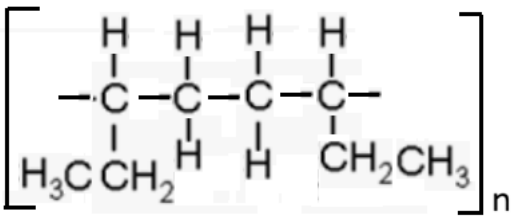
_____*but-1-ene*_____

- Draw the repeating unit of the polymer that will result using this monomer.

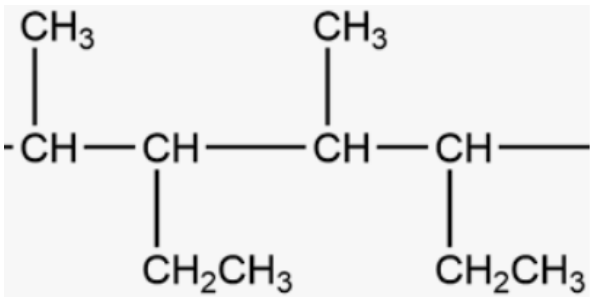
Step 1- Draw the monomers next to each other, as shown below, exposing their carbon-to-carbon double bond(C=C).



Step 1- Break the carbon-to-carbon double bonds(C=C) and join the carbons with single bonds. They now form the polymer carbon backbone

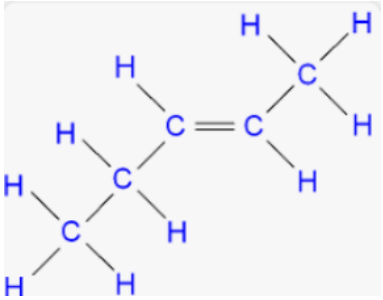


- Consider the repeating unit of the polymer shown below.

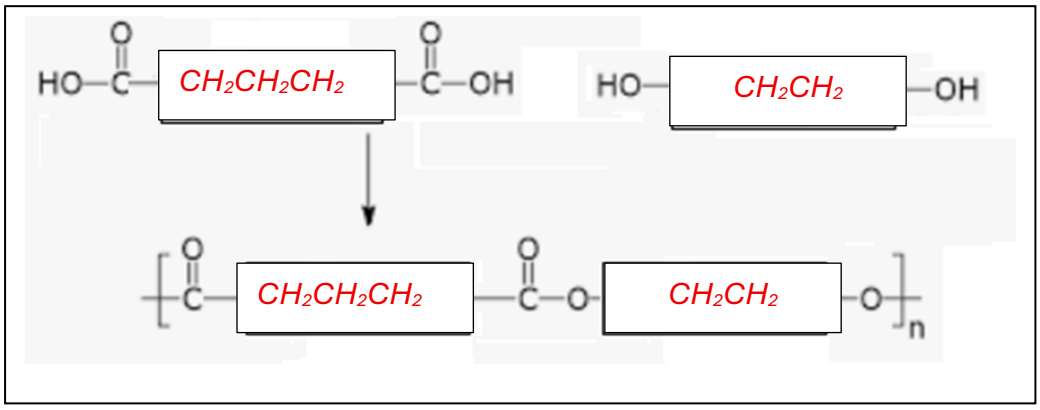


  - Draw its structural formula and write its semistructural formula in the space provided below.
  - Give the systematic name of the monomer shown on the right.

Semistructural _____*pent-2-ene*____*$CH_3CH_2CHCHCH_3$*



Q2 Consider the monomers A and B shown below.



Monomer A is pentanedioic acid and B is ethane-1,2-diol.

i. Write the semistructural formula in the box provided for the polymer repeating unit.

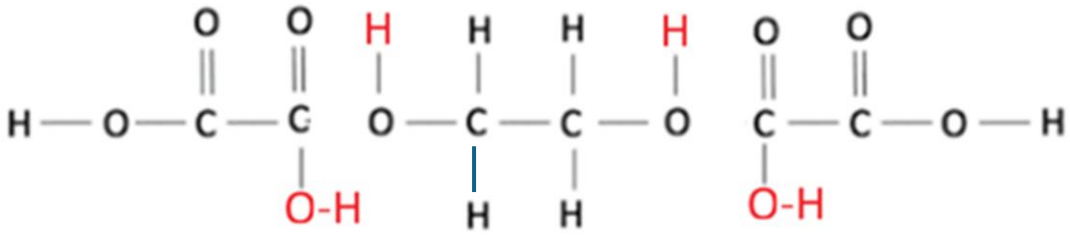
ii. What small molecule is expelled in the formation of the polymer?

_____*H_2O*_____

iii. Name the functional group present in the repeating unit of the polymer.

_____*Ester*_____

Q3. Consider the polymer shown below.



Step 1 – Break the ester bonds between the C-O.

Step 2 – Introduce water back into the bond (hydrolysis). -OH is put back on the C=O and -H back on the O of the alcohol, as shown above.