In Chemistry, "plastic" = ductile + malleable. Plastic is an adjective; a property. "polymer" is a type of material that includes polythene, PVC and Nylon®. Polymers are made from repeating monomer units bonded covalently.

Polymers are made from repeating monomer units bonded covalently.

Branching:
- Low Branching:
  - High-density polythene
  - More Crystalline
  - crystalline structures scatter light, making HDPE opaque
- High Branching:
  - Low-density polythene
  - Amorphous
  - no crystalline structures, so no light scattering, making LDPE transparent

Large Side-Groups:
- Monomer
  - very large
- Polymer
  - Polyethylene: low density

Arrangement of Side-Groups:
- Atactic
  - random arrangement of side-groups
  - less attraction between polymer chains
  - QUITE USELESS (grease)
- Isotactic
  - all side groups face the same direction
  - more attraction between polymer chains
  - USEFUL: clothes, and babies' bottles that don't melt when sterilised

4 structural features that change a thermoplastic polymer's properties (see next section for number 4)

High-density polythene: More Crystalline
Low-density polythene: Amorphous

Branching lowers the density, hardness and melting point of a polymer.

Isotactic polymers can pack more closely and are more dense than atactic polymers.

Polystyrene:
- Atactic
  - random arrangement of side-groups
- Isotactic
  - all side groups face the same direction

USEFUL: clothes, and babies' bottles that don't melt when sterilised

Quite useless (grease)
Mendeleev

1. Made the first modern periodic table
2. Arranged elements by atomic mass
3. Found repeating properties (periods)
4. Predicted many elements
5. He put Co & Ni the other way around!
6. Didn’t find any noble gases

Mendeleev's table had gaps where undiscovered elements would later be placed

Seaborg

Discovered 10 f-block elements during World War II
Practice Questions for 1.1.1a

(i) Mendeleev

1. Elements are said to exhibit periodicity. What do you think this term means?
2. How did Mendeleev organise the elements in his periodic table horizontally?
3. How did Mendeleev organise the elements in his periodic table vertically?
4. Name some of the elements that were not included in Mendeleev’s periodic table.
5. Why were these elements not included?

(ii) Seaborg

1. What is the meaning of the word “transuranium”?
2. How did Seaborg discover ten elements? (Do some quick research to answer this question)

(iii) Modern Periodic Table

1. What are the four blocks of the periodic table called?
2. Name an element in the same group as neon.
3. Name an element in the same period as sodium.
4. Name an element with similar properties to nitrogen.
5. Name a transition metal element.
6. Name an element that is a lanthanide.
The Periodic Table - trends and patterns of properties within The Periodic Table:
atomic number, types of compounds formed, metallic/non-metallic character, chemical reactivity of elements.
Practice Questions for 1.1.1b

(i) Types of Compounds Formed
1. Give five examples of an ionic compound (formula and name).
2. Give five examples of a covalent compound (formula and name).

(ii) Atomic Number (Z)
1. Define atomic number.
2. Why do atomic numbers increase as you go down and right in the periodic table?

(iii) 1st Ionisation Energy
1. What is first ionisation energy?
2. Explain the trend in first ionisation energy, including any exceptions to the rule.

(iv) Metallic Character
1. Define metallic character.
2. What is the most ‘metallic’ element?
3. Describe the trend in metallic character in the periodic table.
4. Explain the trend in metallic character in the periodic table.

(v) Chemical Reactivity
1. Explain the trend in chemical reactivity.
2. Arrange the following reactions in increasing order of energy produced (only reactants shown):
   \[ 2\text{Na} + \text{Cl}_2 \quad 2\text{Na} + \text{F}_2 \quad 2\text{Na} + \text{Br}_2 \quad 2\text{Na} + \text{I}_2 \]
3. Arrange the following reactions in increasing order of energy produced (only reactants shown):
   \[ \text{Cs} + \text{H}_2\text{O} \quad \text{K} + \text{H}_2\text{O} \quad \text{Li} + \text{H}_2\text{O} \quad \text{Na} + \text{H}_2\text{O} \]

(vi) Reason for the Horizontal Trend
1. What is the underlying reason for the horizontal trends in the periodic table?

(vii) Reason for the Vertical Trend
1. What is the underlying reason for the vertical trends in the periodic table?
Dalton, 1808
First to describe atoms in a modern, scientific sense
- Doesn’t explain electricity
+ Idea of “atoms”

Thomson, 1897
Thomson’s Plum Pudding Model
- Doesn’t explain why some of Rutherford’s α-particles bounced back
+ Protons & electrons

Rutherford, 1911
Rutherford shot α-particles through gold foil; some bounced back!
- Why don’t the electrons lose energy and crash into the nucleus?
+ the Nucleus

Bohr, 1913
Basis for our modern atomic model
- Doesn’t explain quantum mechanics
+ Electron Shells

Schrödinger, 1926
Quantum mechanics
- Why are some atoms of the same element heavier?
+ Subshells
+ ‘Shells’ are actually ‘orbitals’

Chadwick, 1932
+ Neutrons!

atomic theory - historical development of the model of atomic theory with contributions from Dalton to Chadwick
(i) General Questions

1. Which scientist is credited for each of the following aspects of the atomic model?
   a) Plum Pudding Model
   b) quantum mechanical model
   c) discovery of neutrons
   d) atoms
   e) nucleus
   f) electrons
   g) electrons orbiting a nucleus
   h) electron shells
   i) electron subshells
   j) orbitals

2. Describe, in detail, Rutherford’s gold foil experiment. What two conclusions did he make?
Limitations of the atomic theory model

Why do electrons move in circular orbits?

Why do shells have particular energies?

Why do some shells allow more than 8 e⁻?

Schrödinger has the answers!

- Electrons can behave like waves
- We can’t know exactly where they are (“cloud of negative charge”)
- It’s called Quantum Mechanics
Practice Questions for 1.1.2b

(i) General Questions

1. Make an acronym to help you remember the three limitations and the three “answers” from Schrödinger.
**Relative Atomic Mass (RAM)**

Weighted mean of all isotopes of an element by their abundances on Earth.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Relative isotopic mass (RAM)</th>
<th>Abundance</th>
<th>Relative Atomic Mass (RAM)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{16}\text{O}$</td>
<td>15.995</td>
<td>99.76%</td>
<td>= 15.957</td>
</tr>
<tr>
<td>$^{17}\text{O}$</td>
<td>16.999</td>
<td>0.04%</td>
<td>= 0.007</td>
</tr>
<tr>
<td>$^{18}\text{O}$</td>
<td>17.999</td>
<td>0.20%</td>
<td>= 0.003</td>
</tr>
</tbody>
</table>

1: Multiply | 2: Add up | 15.967

**Electron Configurations**

- Name of the subshell
- Number of electrons in that subshell
- Each letter-number pair is a subshell

<table>
<thead>
<tr>
<th>Shell</th>
<th>Subshells</th>
</tr>
</thead>
<tbody>
<tr>
<td>First</td>
<td>1s</td>
</tr>
<tr>
<td>Second</td>
<td>2s, 2p</td>
</tr>
<tr>
<td>Third</td>
<td>3s, 3p, 3d</td>
</tr>
<tr>
<td>Fourth</td>
<td>4s, 4p, 4d, 4f</td>
</tr>
<tr>
<td>Fifth</td>
<td>5s, 5p, 5d, 5f</td>
</tr>
<tr>
<td>Sixth</td>
<td>6s, 6p, 6d, 6f</td>
</tr>
</tbody>
</table>

**Aufbau Principle**

Means “building up”

Helps you to remember the order in which the subshells are filled.

**Drawing Electron Configurations**

- Max. 2 electrons per atomic orbital
- Place 1 electron in each atomic orbital before placing 2 in the 1st atomic orbital
- 3d⁴ and 3d⁵ are unstable. Cu & Cr ‘steal’ an electron from 4s instead.

Cu: $4s^1 3d^5$
Cr: $4s^1 3d^{10}$

**Atomic Theory**

- Mass number, isotopes, calculation of relative atomic mass, electron configuration using subshells
Practice Questions for 1.1.2c

(i) Nuclide Representation

1. Give the nuclide representations for the following elements:
   a) an atom with 6 protons, 6 electrons and 6 neutrons
   b) an atom with 7 protons, 7 electrons and 7 neutrons
   c) an atom with 8 protons, 8 electrons and 10 neutrons
   d) an atom with 13 protons, 10 electrons and 14 neutrons
   e) an atom with 1 proton, no electrons and no neutrons
   f) an atom with 17 protons, 18 electrons and 18 neutrons
   g) an atom with 17 protons, 18 electrons and 20 neutrons

(ii) Isotopes

1. Define an isotope.
2. Which of the atoms in part (i) are isotopes of each other?

(ii) RIM & RAM

1. Define RIM & RAM and give their units.
2. Calculate the RAM of chlorine using the following data. 75% $^{35}$Cl and 25% $^{37}$Cl

(iv) Electron Configurations

1. Give the full electron configurations of the following:
   a) a hydrogen atom
   b) a nitrogen atom
   c) an argon atom
   d) a calcium atom
   e) an excited calcium atom
   f) a Ca$^{2+}$ ion

(v) Drawing Electron Configurations

1. Draw the electron configurations of all the species in part (iv).


6.02 \times 10^{23} = \text{“Avogadro’s constant”} = N_A = 1 \text{ mole}

**Definition**

1 mole is the no. of atoms in exactly 12 g of $^{12}$C

### Mole Formula 1

- **no. of moles** (mol)
- **mass (g)**
  
  \[ n = \frac{m}{M} \]
  
  molar mass (g/mol)

### Mole Formula 2

- **no. of moles** (mol)
  
  \[ N = n \times N_A \]
  
  \[ 6.02 \times 10^{23} \]

**Empirical formula:**

- simplest whole-number ratio

- \( \text{CH}_3, \text{CH}_2\text{O}, \text{H}_2\text{O}, \text{CO}_2 \)

**Molecular formula:**

- actual number of atoms in the molecule

- \( \text{C}_2\text{H}_6, \text{C}_6\text{H}_{12}\text{O}_6, \text{H}_2\text{O}, \text{CO}_2 \)

**To find an empirical formula from mass composition:**

<table>
<thead>
<tr>
<th></th>
<th>C</th>
<th>H</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>m</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>n</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ratio</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Extension:** multiply this ratio by \( \frac{M_{\text{molecule}}}{M_{\text{empirical}}} \) to get the molecular formula

---

VCEasy VISUAL CHEMISTRY 1.1.3

the mole concept including empirical and molecular formulas, percentage composition, Avogadro’s constant
Practice Questions for 1.1.3

(i) Mole

1. What is a mole?

(ii) Mole formula 1

1. Find the number of moles of molecules in the following samples:
   a) 18 grams of water
   b) 1 kilogram of silicon dioxide
   c) 5 grams of N₂O₄
   d) 200 grams of P₄

(ii) Mole formula 2

1. Write the following in moles (in scientific notation):
   a) 6.02×10²² moles of protons
   b) 6.02×10²³ calcium ions
   c) 4.1×10¹⁹ water molecules
   d) one million dollars

(iv) EF & MF

1. What are the empirical formulae of BF₃, O₂, C₁₀H₂₀ and Fe₂O₃?

(v) Percentage Composition

1. Find the empirical formulae of the following compounds:
   a) 60.0% carbon, 13.4% hydrogen, 26.6% oxygen
   b) 37.5% carbon, 12.6% hydrogen, 49.9% oxygen
   c) 14.1% carbon, 2.4% hydrogen, 83.5% chlorine

2. Find the molecular formulae of the compounds in question 1 if their relative molecular masses are 60.1, 32 and 84.9 respectively.
### Interpretation of Data from Mass Spectrometry

**Mass-to-charge (m/z) ratio**

- **Relative abundance (R.A.) %**
  - Sometimes called “relative peak area”

**Abundances**

1. **Calculate abundances**
   
   \[ \text{abundance} = \text{relative abundance} \div \Sigma (\text{relative abundances}) \]

2. **Find weighted average**
   
   \[ A_r = \Sigma (\text{relative atomic mass} \times \text{abundance}) \]

   \[ A_r = (26 \times 0.79) + (27 \times 0.09) + (28 \times 0.11) + (30 \times 0.01) \]

   \[ A_r = 26.35 \]

   If working backwards, you may need to substitute for \( x \) and \( (1-x) \)

**Example**

- 26: 100% (100/100)
- 27: 11% (11/100)
- 28: 14% (14/100)
- 29: 1% (1/100)
- 30: 0% (0/100)

If the values aren’t percentages, (i.e. they don’t add up to 100), then you will need to convert them into percentages first.
Practice Questions for 1.1.4

(i) Interpretation

1. Draw the mass spectrum graph you would see if you analysed the following samples:
   a) a sample of aluminium containing 79% $^{26}\text{Al}$, 9% $^{27}\text{Al}$, 11% $^{28}\text{Al}$ and 1% $^{30}\text{Al}$
   b) a sample of carbon containing 98.89% $^{12}\text{C}$ and 1.11% $^{13}\text{C}$
   c) a sample of boron containing 80.0% $^{11}\text{B}$ and 20.0% $^{10}\text{B}$

2. Calculate the relative atomic masses (RAM) for each of the above samples

3. Calculate the percentage abundance of each isotope for the following elements:
   a) boron, which as two isotopes: $^{11}\text{B}$ (relative isotopic mass = 11.01) and $^{10}\text{B}$ (relative isotopic mass = 10.01). The relative atomic mass of boron is 10.81.
   b) gallium, which has two isotopes: $^{69}\text{Ga}$ (relative isotopic mass = 68.93) and $^{71}\text{Ga}$ (relative isotopic mass = 70.92). The relative atomic mass of gallium is 69.72.
   c) silver, which as two isotopes: $^{107}\text{Ag}$ (relative isotopic mass = 106.91) and $^{109}\text{Ag}$ (relative isotopic mass = 108.91). The relative atomic mass of silver is 107.90.
metallic bonding

- Very Strong
- Neither Covalent nor Ionic
- Only Occurs in Metals

mobile, delocalised valence electrons

cell

conduct heat

conduct electricity

lustrous

dense

opaque

delts = Ductile + Malleable

HIGH MELTING POINT

HIGH TENSILE STRENGTH

HARD (RESISTS SCRATCHING)

crystals have regular, geometric shapes

metallic bonding

THIS IS A METAL “LATTICE”

4 Ways to Modify Metals

Heat Treatment

Annealing

Quenching

Tempering

Soft

Hard

HARD, LESS BRITTLE

Alloys

Substitutional Alloys

Interstitial Alloys

Work Hardening

knocks crystals into a tighter conformation results in harder, more brittle metal

Surface Coating

coat the metal with another metal e.g. galvanisation

layers of cations in the metal lattice don’t slide over each other so easily in an alloy

models of bonding to explain observed properties including melting temperature, electrical conductivity, chemical reactivity, shape, polarity of bonds, intermolecular forces - metals
Practice Questions for 1.2.1a

(i) Metallic Bonding

1. Draw a metallic lattice of sodium.
2. Draw a metallic lattice of magnesium.
3. Draw a metallic lattice of zinc.
4. Explain why the cations in a metal lattice do not repel each other.
5. Explain why copper feels colder than wood, even when both are at the same temperature.
6. Make up an acronym to help you remember all ten properties of metals.
7. For each of the ten properties, think of one example metal and write how that property is useful.

(ii) 4 Ways to Modify Metals

1. Using words and pictures, describe each of the following metal treatments:
   a) heat treatments (three types)
   b) alloys (two types)
   c) work hardening
   d) surface coating

b) Give one example of each of the metal modification methods in question 1 and state why the properties of those metals are useful.
Ionic compounds \( \Delta \text{EN} > 1.7 \)

- High MP & BP
- Hard & brittle
- Conduct electricity **ONLY** when liquid or aqueous

**Strong ionic bonds:**
- Brittle crystals with a high MP
- Doesn’t conduct electricity as a solid: there are no free-moving charged particles
- Conduct as a liquid or in aqueous solution: there are free-moving charged particles
- Forms a **lattice** to maximise the forces of attraction

**Ionic compounds are brittle**

When knocked out of place, the cations and anions can be knocked into a configuration where they repel each other, causing a **fracture**.

**Ionic bonds involve transfer of electrons from one atom to another**

models of bonding to explain observed properties including melting temperature, electrical conductivity, chemical reactivity, shape, polarity of bonds, intermolecular forces - ionic compounds
Practice Questions for 1.2.1b

(i) Ionic Lattices

1. Why do ionic lattices have such high melting points and boiling points?
2. Why do ionic lattices not conduct electricity?
3. Why do ionic compounds conduct electricity when they are aqueous or molten?

(ii) Brittle

1. Explain using words and diagrams why large crystals of sodium chloride can be broken with a hammer.

(iii) Transfer of Electrons

1. Draw electron transfer diagrams for the following reactions. (Only draw the outer shells.)
   a) lithium and chlorine to produce lithium chloride.
   b) aluminium and fluorine to produce aluminium fluoride.
   c) calcium and nitrogen to make calcium nitride.
   d) magnesium and sulphur to make magnesium sulphide.
Covalent Bonding \( \Delta EN \leq 1.7 \)

Covalent molecules have low MP & BP
Covalent network lattices don’t conduct electricity; high MP & BP
Covalent layer lattices conduct electricity; high MP & BP

Allotropes of Carbon

- **Diamond**
  - High BP due to strong intermolecular forces
  - Each C is bonded covalently to 4 other carbon atoms.
  - \(-78.5^\circ\text{C}\)

- **Graphite**
  - Low BP due to weak intermolecular forces
  - Each C is bonded covalently to 3 others. The 4th \(s^2\) is delocalised and conducts electricity.
  - \(-182^\circ\text{C}\)

- **Coal**
  - Each C is bonded to 4 others. Bonding is “amorphous”.
  - \(-3600^\circ\text{C}\)

- **Nanotubes**
  - Strongest known material.
  - >3500°C

Compounds of Carbon

- There are thousands, but here are just two examples:
  - \(\text{O} = \text{C} = \text{O}\)

- **Solid up to**
  - High BP due to strong intermolecular forces
  - \(-78.5^\circ\text{C}\)

- **Solid up to**
  - Low BP due to weak intermolecular forces
  - \(-182^\circ\text{C}\)

Shapes of Covalent Molecules (VSEPR)

- **Tetrahedral**
  - 0.5 Lone Pairs
  - LEWIS STRUCTURE
  - ELECTRON DOTS

- **Triangular Pyramid**
  - 2 Lone Pairs
  - LINEAR

- **Bent or “V-Shaped”**
  - 3 Lone Pairs
  - LINEAR

- **Linear**
  - 2 Lone Pairs (Zero Lone Pairs)

Net Dipole

- Sum of all the polarities of all the polar bonds in the molecule. Add the up like vectors.

\[ \Delta EN: \quad 0.0 \rightarrow 0.5 \rightarrow 1.7 \]

the ionic character of a bond increases as \( \Delta EN \) increases

- **Intramolecular Bonds**
  - within a molecule
  - C—H: non-polar covalent
  - dispersion forces: **WEAK**
  - insoluble
  - very small increase

- **Intermolecular Bonds**
  - between two molecules
  - O—H: polar covalent
  - dipole-dipole interactions or “hydrogen bonding”
  - soluble
  - medium increase

- **Solubility in Water**
  - **Insoluble**

- **Effect on BP & MP**
  - **Soluble** (ions dissociate)
  - very large increase

models of bonding to explain observed properties including melting temperature, electrical conductivity, chemical reactivity, shape, polarity of bonds, intermolecular forces - molecular substances, network lattices, layer lattices
Practice Questions for 1.2.1c

(i) Allotropes

1. Compare and contrast diamond, graphite and coal in a table.

2. For each of the five allotropes of carbon on the previous page, give one example of how they are used. Explain why their unique properties are useful in each case.

3. Why do all allotropes of carbon have such high boiling points?

(ii) Compounds of Carbon

1. Why do CO$_2$ and CH$_4$ have such low boiling points?

(iii) 3 Ways to Draw a Molecule

1. Draw the electron dots, Lewis structures and valence structures of the following covalent molecules: CH$_4$, SiBr$_4$, HCl, CH$_3$OH, H$_2$O, H$_2$S, NCl$_3$, PH$_3$, PH$_2$Cl, I$_2$, BeH$_2$, H$_2$Te, CF$_4$. Arrange them in a table.

(iv) Shapes of Covalent Molecules (VSEPR Theory)

1. Explain the underlying principle of VSEPR theory.

2. State the shapes of each of the molecules you drew in part (iii).

(v) Electronegativity Differences ($\Delta$EN)

1. Describe the following intramolecular bonds: non-polar covalent bonds, polar covalent bonds, ionic bonds.

2. Describe the following intermolecular bonds: dispersion forces, dipole-dipole interactions, hydrogen bonds, ionic-ionic interactions

3. Arrange all the bonds from questions 1 and 2 in increasing order of strength.

4. Arrange the following molecules in order of boiling point: CH$_4$, CH$_2$Cl$_2$, CH$_3$Cl.

5. Draw $\partial^+$ and $\partial^-$ on the following bonds: O—H, F—H, P—Cl, C—H, C—Cl, C—F, N—H, C=O

6. Draw $\partial^+$ and $\partial^-$ on the molecules in part (iii).

7. Explain, using words and pictures, what an “ion-dipole interaction” is.

(vi) Net Dipole

1. Using an arrow, draw the net dipole on each of the molecules in part (iii) (if any).
Limitations of the 3 bonding models
(metallic, ionic and covalent)

**Metallic bonding model limitations**
- Melting Point \( (\text{Hg} \ll \text{W}) \)
- Density \( (\text{Li} \ll \text{Os}) \)
- Conductivity \( (\text{Fe} \ll \text{Cu}) \)

**Ionic bonding model limitations**
- Shape (cubic, needle, etc.)
- Solubility \( (\text{AgCl} \ll \text{NaCl}) \)

**Covalent bonding model limitations**
- Different number of atoms in each molecule of a Group 3 element
  \( \text{O}_2 \ll \text{P}_4 \ll \text{S}_8 \)
Practice Questions for 1.2.2

(i) Metallic Bonding
1. What are the limitations of this bonding model?

(ii) Ionic Bonding
1. What are the limitations of this bonding model?

(iii) Covalent Bonding
1. What are the limitations of this bonding model?
A carbon backbone is the longest –C–C– chain in a molecule.

An organic compound is a compound made mostly from C and H. Some organic compounds also have functional groups attached (such as –OH or –NH₂).

A homologous series is a series of organic compounds with similar chemical properties (same functional group) and in which each member differs by a –CH₂– group from the previous member.

### Backbone

<table>
<thead>
<tr>
<th>Carbon</th>
<th>Saturated (alkanes)</th>
<th>Unsaturated (alkenes)</th>
</tr>
</thead>
<tbody>
<tr>
<td>meth– 1 carbon</td>
<td>methane</td>
<td>CH₄</td>
</tr>
<tr>
<td>eth– 2 carbons</td>
<td>ethane</td>
<td>CH₃CH₃</td>
</tr>
<tr>
<td>prop– 3 carbons</td>
<td>propane</td>
<td>CH₃CH₂CH₃</td>
</tr>
<tr>
<td>but– 4 carbons</td>
<td>butane</td>
<td>CH₃CH₂CH₂CH₃</td>
</tr>
<tr>
<td>pent– 5 carbons</td>
<td>pentane</td>
<td>CH₃CH₂CH₂CH₂CH₃</td>
</tr>
<tr>
<td>hex– 6 carbons</td>
<td>hexane</td>
<td>CH₃CH₂CH₂CH₂CH₂CH₃</td>
</tr>
</tbody>
</table>

### Properties

<table>
<thead>
<tr>
<th>Alkanes</th>
<th>Alkenes</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Combustion</td>
<td>1. Add bromine solution</td>
</tr>
<tr>
<td>Alkane + O₂ (g) → CO₂(g) + H₂O(g)</td>
<td>Colourless</td>
</tr>
<tr>
<td>C H O</td>
<td>Colourless</td>
</tr>
</tbody>
</table>

### Properties

<table>
<thead>
<tr>
<th>Alkanes</th>
<th>Alkenes</th>
</tr>
</thead>
<tbody>
<tr>
<td>2. Add hydrogen gas makes an alkane</td>
<td>2. Add hydrogen gas</td>
</tr>
<tr>
<td>3. Add steam makes an alcohol</td>
<td>Colourless</td>
</tr>
<tr>
<td>4. Addition polymerisation (see Unit 1.2.6)</td>
<td></td>
</tr>
</tbody>
</table>

### Full Structural Formulae

- Pentane
- 2-Methylbutane
- 2,2-Dimethylpropane

### Properties

- Combustion
- Balance the equation in this order

### Properties

- Add bromine solution
- Add hydrogen gas
- Add steam
- Addition polymerisation (see Unit 1.2.6)
Practice Questions for 1.2.3

(i) Naming

1. What is a carbon backbone?

2. Draw a compound with a five-carbon backbone. (Draw the full structural formula using the correct bond angles.)

(ii) Prefixes & Suffixes

1. How many carbon atoms are in the following molecules?
   a) propane
   b) pentane
   c) methane
   d) methylpropane
   e) 3-methylhexane
   f) but-1-ene
   g) but-2-ene
   h) 3,3-dimethylpentane

2. Draw the full structural and semi-structural formulae of all the molecules in question 1.

(iii) Properties of Alkanes

1. Think of ten hydrocarbon-based fuels. Write balanced equation of complete combustion for these ten fuels.

(iv) Properties of Alkenes

1. What are the four reactions of alkenes?

2. Which of these four reactions involves:
   a) decolourisation?
   b) a high temperature?
   c) a catalyst?
   d) production of an alcohol?
   e) making products with a very high molar mass?
Structural Isomers

Compounds with the same molecular formula but different structural formulae.

![Structural Isomers Diagram]

**Butane**

\[
\text{C}_4\text{H}_{10}
\]

**2-Methylpropane**

\[
\text{C}_4\text{H}_{10}
\]

Structural isomers of \(\text{C}_4\text{H}_{10}\)
Practice Questions for 1.2.4

(i) Structural Isomers

1. What is a structural isomer?
2. Draw all the structural isomers of butane (C₄H₁₀).
3. Draw all the structural isomers of pentane (C₅H₁₂).
4. Draw all the structural isomers of hexane (C₆H₁₄).
Surface Tension

Liquid molecules are repelled from the surface because they only experience downward intermolecular forces.

Water-glass interaction is stronger than water-water interaction due to water ‘wets’ glass. Therefore, water forms droplets on polythene. Polythene has a lower surface energy than water.

Glass has a higher surface energy than water.

Nanoparticles

Tiny particles nanometers in size

1. Very high surface area-to-volume ratio leads to better catalysts & increased friction (e.g. gecko-gloves).

2. Electrons aren’t 100% delocalised leads to more heat sensitivity & colours! (e.g. nano-gold = ruby gold).

Stainless clothing

Transparent sunscreen

UV-opaque glass

Designer cancer drugs

Four uses of nanoparticles:

- Stainless clothing
- Transparent sunscreen
- UV-opaque glass
- Designer cancer drugs

Behaviour of surfaces and the application of surface chemistry in nanotechnology.
Practice Questions for 1.2.5

(i) Surface Tension

(ii) Nanoparticles

(iii) 4 Uses of Nanoparticles
In Chemistry, ‘plastic’ = ductile + malleable. Plastic is an adjective; a property. ‘polymer’ is a type of material that includes polythene, PVC and Nylon®. Polymers are made from repeating monomer units bonded covalently.

**Branching**
- **Low Branching**
  - High-density polyethene
  - More Crystalline
  - Crystalline structures scatter light, making HDPE opaque

- **High Branching**
  - Low-density polyethene
  - Amorphous
  - No crystalline structures, so no light scattering, making LDPE transparent

**Large Side-Groups**
- Illustrated by polystyrene
- Large side-groups prevent dense packing and lowers the density of a polymer.

**Arrangement of Side-Groups**
- Illustrated by polypropene
- Isotactic polymers can pack more closely and are more dense than atactic polymers.
  - Atactic
    - Random arrangement of side-groups
    - Less attraction between polymer chains
    - QUITE USELESS (grease)
  - Isotactic
    - All side groups face the same direction
    - More attraction between polymer chains
    - USEFUL: clothes, and babies' bottles that don't melt when sterilised

Addition polymers - relationship between structure, properties and applications
Practice Questions for 1.2.6a

(i) Definitions

1. Define the following terms:
   a) monomer
   b) polymer
   c) plastic
2. What’s the difference between “polymer” and “plastic”?

(ii) Branching

1. What is the difference between “amorphous” and “crystalline” polymers?
2. What is the difference between HDPE and LDPE?
3. Which polymer has more branching: HDPE or LDPE?
4. How does branching change the properties of a polymer?
5. Why is cling film, which is made from LDPE, transparent?

(iii) Large Side-Groups

1. What is the effect on a polymer of having large side-groups in its monomer units?
2. Give an example of a polymer with large side-groups on its monomer units.

(iv) Arrangement of Side-Groups

1. Carefully-controlled reaction conditions can change whether atactic or isotactic polymers are produced.
   a) Describe the differences in structure between atactic and isotactic polymers.
   b) Describe the differences in properties between atactic and isotactic polymers.
**Thermoplastic polymers** have no cross-links.

- e.g. 3D printer material

**Elastomers** have occasional cross-links or hydrogen bonding between side-groups in the chains

**Thermosetting polymers** have extensive cross-links and do not melt when heated. They decompose instead.

Cross-links are made of **sulphur**
- cross-links are strong, covalent bonds
- created through a process known as **vulcanisation**
- more sulphur, more cross-links, more elastic memory

- chains of rubber polymer before cross-linking
- sulfur, heat (vulcanisation)
- Hydrogens not shown

**To make a Thermoplastic from Monomers:**

\[ nC_2H_4(g) \rightarrow (C_2H_4)_n(s) \]

**Monomer** | **Polymer**
---|---

- addition polymers - synthesis, cross-linking
Practice Questions for 1.2.6b

(i) Degree of Cross-Linking

1. Compare thermoplastic polymers, elastomers and thermosetting polymers.
2. What happens when you heat thermoplastic polymers, elastomers and thermosetting polymers?
3. Describe cross-links in polymers.
4. Industries sometimes alter the reaction conditions and reaction reagents so that the polymers they are producing have cross-links. What is the role of cross-links in polymers?
5. Describe, in terms of intermolecular bonds, why gaseous monomer units can form a solid polymer.
Polymers can be customised in four ways:

1. Using two or more different monomers

2. Changing the side-groups

3. Rearranging the side-groups (atactic/isotactic)

4. Adding plasticisers

Customised polymers can have enhanced properties, such as:

- flame retardance
- resistance to UV-light degradation
- waterproofness
- flexibility
- impact resistance
- transparency

Plasticisers sit between the polymer chains, separating them slightly. They reduce the effect of intermolecular forces, making the polymer softer and more flexible.
Practice Questions for 1.2.6c

(i) Customised Polymers

1. Explain each of the four ways in which polymers can be customised.
2. Think of an acronym to help you remember these four customisations.
Hydrogen Bonding gives water its very high boiling point

Theoretical # of intermolecular bonds formed

- 4 per molecule for H₂O in the absence of hydrogen bonding (1.24)
- 2 per molecule for HF (1.78)
- 2 per molecule for NH₃ (0.84)

Therefore...

- H₂O has a very high boiling point
- HF and NH₃ have lower boiling points

Hydrogen Bonding gives water its very high latent heat values

- The latent heat of fusion of water is the amount of energy needed to change a fixed amount of water from a solid to a liquid at 0°C.
- The latent heat of vaporisation of water is the amount of energy needed to change a fixed amount of water from a liquid to a gas at 100°C.
- These values are high because of the strong hydrogen bonds that must be broken during each change of state.

Water has a high specific heat capacity

- It requires a lot of energy to heat water compared with other substances.

\[
E = \text{SHC} \times m \times \Delta T
\]

Ice is less dense than water

- Below 0°C, all four hydrogen bonds stabilise into a crystal structure (ice). In ice, the molecules are spaced more widely apart than in liquid water.
Practice Questions for 2.1.1a

(i) Introduction to Water
1. List five to ten roles of water in the environment.
2. Write the balanced photosynthesis equation.
3. Write the balanced equation for the complete combustion of glucose in oxygen.

(ii) High Boiling Point
1. What is hydrogen bonding?
2. Draw two water molecules and draw a hydrogen bond between them.
3. Explain how hydrogen bonding gives water an unusually high boiling point.
4. Why is the boiling point of H\textsubscript{2}S so much lower than the boiling point of H\textsubscript{2}O?
5. Why is the boiling point of H\textsubscript{2}O so much higher than the boiling point of NH\textsubscript{3}?

(iii) High Latent Heat Values
1. What are the two types of latent heat values?
2. Explain these two types of latent heat values in simple terms.
3. Why does ice require so much energy to melt it?
4. Why does water require so much energy to boil it?

(iv) High Specific Heat Capacity
1. What is a specific heat capacity?
2. Water has a very high specific heat capacity. Name three materials with very low specific heat capacities.
3. Why does water require so much energy to warm it up?
4. Why does water cool down so slowly?

(v) Ice is Less Dense than Water
1. Explain, in terms of intermolecular bonding, why ice is less dense than water.
1. **Particles of solute are separated from one another.**
2. **Particles of solvent are separated from one another.**
3. The solute and solvent particles attract each other.

**3 types of substances in water**

### 1. Ionic Substances in Solution

Example: NaCl(s) → Na⁺(aq) + Cl⁻(aq)

![Diagram showing the water molecules surrounding the ionic lattice and pulling the anions and cations into the surrounding solution.](image)

### 2. Polar Molecules in Solution that ionise

Example: H₂O(l) + HCl (g) → H₃O⁺(aq) + Cl⁻(aq)

![Diagram showing the reaction between water and HCl, forming hydronium ions.](image)

The solute loses a proton and ionises. That proton bonds with a water molecule to form a hydronium ion (H₃O⁺).

### 3. Polar Molecules in Solution that don’t ionise

Example: C₂H₅OH(l) → C₂H₅OH(aq)

![Diagram showing the reaction between ethanol and water molecules.](image)

When ethanol is added to water, hydrogen bonds form between the ethanol and water molecules and the ethanol dissolves.

Water and ethanol molecules mix freely with each other, held together in the solution by hydrogen bonds.

---

Role of water in maintaining life in the environment - ways in which substances behave in water: the dissociation of soluble ionic solutes; the ionisation of polar molecules such as acids; the separation of non-ionising polar molecules such as ethanol.
Practice Questions for 2.1.1b

(i) When a Substance Dissolves in Water
1. What happens when a substance dissolves in water?

(ii) Ionic Substances in Solution
1. Define an “ionic” substance.
2. Explain, using words and pictures, what happens when sodium chloride dissolves in water.
3. When ionic substances dissolve, how is hydrogen bonding involved?

(iii) Polar Substances that Ionise in Solution
1. Define “ionisation”.
2. What is a “hydronium ion”?
3. Explain, using words and pictures, what happens when hydrochloric acid dissolves in water.
4. When polar substances that ionise in solution are dissolved in water, how is hydrogen bonding involved?

(iv) Polar Substances that Don’t Ionise in Solution
1. Explain, using words and pictures, what happens when ethanol dissolves in water.
2. When polar substances that don’t ionise in solution are dissolved in water, how is hydrogen bonding involved?
Solubility
- refers to the **maximum mass** of a substance that can dissolve in a given quantity of solvent at a given temperature
- solute that can no longer remain in solution (e.g. because the temperature has decreased) will crystallise or **supersaturate**. Heat packs rely on the sudden recrystallisation of supersaturated solute, which is an exothermic process.
- temperature and pressure both increase solubility of solids in liquids
- temperature decreases solubility of gases in liquids (e.g. O₂ in water)

**Solution Formula**

\[ n = c \times V \]

\[ \text{mol} = \text{mol} \cdot \text{L}^{-1} \times \text{L} \]

to find concentration

**Units of Concentration**

<table>
<thead>
<tr>
<th>Units</th>
<th>ppm</th>
<th>ppm</th>
<th>mg L⁻¹</th>
<th>mg L⁻¹</th>
</tr>
</thead>
<tbody>
<tr>
<td>ppb</td>
<td>10⁶</td>
<td>1</td>
<td>1 L⁻¹</td>
<td>1 L⁻¹</td>
</tr>
<tr>
<td>µg g⁻¹</td>
<td>10³</td>
<td>10⁴</td>
<td>1 L⁻¹</td>
<td>1 L⁻¹</td>
</tr>
<tr>
<td>%</td>
<td>100</td>
<td>1</td>
<td>1 L⁻¹</td>
<td>1 L⁻¹</td>
</tr>
<tr>
<td>% mass</td>
<td>100</td>
<td>1</td>
<td>1 L⁻¹</td>
<td>1 L⁻¹</td>
</tr>
</tbody>
</table>

**Dilution Formula**

\[ C_1V_1 = C_2V_2 \]

\[ \text{mol L}^{-1} \cdot \text{L} = \text{mol L}^{-1} \cdot \text{L} \]

when adding distilled water to a sample

**Mixtures Formula**

\[ C_1V_2 + C_2V_2 = C_3(V_1+V_2) \]

to find concentration when mixing two solutions together

**Precipitation** occurs when amount of solute exceeds the solubility limit at that temperature

\[ \text{Pb(NO}_3\text{)}_2(aq) + 2\text{KI}(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq) \]

Remove the spectator ions to get the simplified equation:

\[ \text{Pb}^{2+}(aq) + 2\text{I}^{-}(aq) \rightarrow \text{PbI}_2(s) \]

when mixing aqueous solutions, ions can exchange in a double displacement reaction. Any resulting solids are precipitates.

**Water purification** is a four-step process

- **Flocculation**
- **Settling of the ‘floc’**
- **Filtering**
- **Chlorination**

**Hard water** is caused by Ca²⁺ ions. Ca²⁺ binds to the stearate of sodium stearate (active ingredient of soap) to produce calcium stearate, which is **scum**.

Add **LIME & ALUM** to waste water

- **LIME**
  \[ \text{Ca(OH)}_2(s) \rightarrow \text{Ca}^{2+}(aq) + 2\text{OH}^{-}(aq) \]
- **ALUM**
  \[ \text{Al}^{3+}(aq) + 3\text{OH}^{-}(aq) \rightarrow \text{Al(OH)}_3(s) \]
  jelly-like, traps fine particles and removes colour from the water

through sand & gravel

add chlorine gas

\[ \text{Cl}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HOCl}(aq) + \text{H}^+(aq) + \text{Cl}^{-}(aq) \]
kills bacteria

role of water in maintaining life in the environment - maintaining water quality: solubility, precipitation reactions, pH
Practice Questions for 2.1.1c

(i) **Solubility**
1. Define the following terms: solute, solvent, solubility, crystallisation, supersaturation.
2. What will happen to the solubility of sodium chloride in water if the temperature is increased?
3. What will happen to the solubility of water vapour in air if the pressure is decreased?
4. When water boils, what is in the bubbles that form?

(ii) **Solution Formula**
1. Find the number of moles of solute in each of the following solutions:
   a) 0.200 L of 0.520 M solution
   b) 0.568 L of 0.110 M solution
2. Find the concentration (in mol L\(^{-1}\)) of the following solutions:
   a) 1.0 mole of solute dissolved in 2.8 litres of solvent
   b) 40 g of sodium chloride dissolved in 310 mL litre of water
   c) 0.01 mol of sodium chloride dissolved in mol of water

(iii) **Dilution Formula**
1. Find the concentration after 380 mL of solvent is added to each of the solutions in part (ii) question 2.

(iv) **Mixtures formula**
1. The solutions in question 1a and 1b in the “solution formula” section are mixed. What is the final volume and final concentration?
2. 410 mL of 1.0 M NaCl solution and 319 mL of 0.8 M NaCl solution are mixed. What is the final volume and concentration?
3. 92 mL of 0.80 M KOH solution and 155 mL of 1.5 M KOH solution are mixed. What is the final volume and concentration?

(v) **Precipitation**
1. What is a “spectator ion”?
2. Write full and ionic equations to depict what happens when zinc metal is added to copper (II) sulphate solution, producing copper metal and zinc sulphate solution.
3. Write full and ionic equations to depict what happens when silver nitrate solution is added to sodium chloride solution, producing sodium nitrate solution and a silver chloride precipitate.

(vi) **Water Purification**

(vii) **Hard Water**
3 Desalination Methods

D Distillation  
- boil seawater then condense the steam

R Reverse Osmosis  
- force seawater through a selectively permeable membrane

I Ion Exchange  
- seawater travels through a cation-exchange resin, then through an anion-exchange resin

role of water in maintaining life in the environment - desalination, including the principles of distillation
Practice Questions for 2.1.1d

(i) Distillation

(ii) Reverse Osmosis

(iii) Ion Exchange
Acids & Bases

Indicators change colour depending on the pH of their environment

Brønsted–Lowry definitions

acids are proton donors, pH < 7
bases are proton acceptors, pH > 7

6 Reactions of Acids

1. Acid + Metal → salt + H₂
2. Acid + Metal Hydroxide → salt + H₂O
3. Acid + Metal Oxide → salt + H₂O
4. Acid + Metal Carbonate → salt + H₂O + CO₂
5. Acid + Metal Hydrogen Carbonate → salt + H₂O + CO₂
6. Acidic non-metal oxide + base → salt + H₂O

Amphiprotic substances can donate or accept a proton can be acids or bases

<table>
<thead>
<tr>
<th>Donates a proton to form:</th>
<th>Accepts a proton to form:</th>
</tr>
</thead>
<tbody>
<tr>
<td>OH⁻, CO₂⁻, HPO₄²⁻, PO₄³⁻, SO₄²⁻</td>
<td>H₂O, HCO₃⁻, H₂CO₃, H₃PO₄, H₂PO₄⁻, H₂SO₄⁻</td>
</tr>
</tbody>
</table>

Strong acids dissociate all their protons

<table>
<thead>
<tr>
<th>Strong acid</th>
<th>Weak acid</th>
<th>Weak base</th>
<th>Strong base</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl(g) + H₂O(l) → H₃O⁺(aq) + Cl⁻(aq)</td>
<td>CH₃COOH(l) + H₂O(l) → CH₃COO⁻(aq) + H₃O⁺(aq)</td>
<td>NH₄⁺(aq) + H₂O(l) → NH₃(aq) + H₃O⁺(aq)</td>
<td>O²⁻(aq) + H₂O(l) → OH⁻(aq) + OH⁻(aq)</td>
</tr>
</tbody>
</table>

Polyprotic acids can donate >1 proton. For example:

1. H₃PO₄(aq) + H₂O(l) → H₂PO₄⁻(aq) + H₃O⁺(aq)
2. H₂PO₄⁻(aq) + H₂O(l) → HPO₄²⁻(aq) + H₃O⁺(aq)
3. HPO₄²⁻(aq) + H₂O(l) → PO₄³⁻(aq) + H₃O⁺(aq)

acids and bases: proton transfer; common reactions of acids; strong and weak acids and bases; polyprotic acids; amphiprotic substances
Practice Questions for 2.1.2

(i) Brønsted-Lowry Theory

1. What is Brønsted-Lowry theory?

(ii) Conjugate Acids & Conjugate Bases

1. Define the following terms: Brønsted-Lowry acid, Brønsted-Lowry base, conjugate acid and conjugate base.

2. For each of these reactions, label the Brønsted-Lowry acid, the Brønsted-Lowry base, the conjugate acid and the conjugate base. Remember to include all states.
   a) ethanoic acid (CH₃COOH) dissolving in water
   b) nitric acid (HNO₃) dissolving in water
   c) sulphuric acid (H₂SO₄) dissolving in water

(iii) 6 Reactions of Acids

1. Illustrate the six reactions of acids in your notebooks. Add explanations and word equations to each one.

(iv) Amphiprotic Substances

1. Using water as an example, explain the meaning of “amphiprotic”.

(v) Strong and Weak Acids and Bases

1. What is the difference between strong and weak acids?

2. Why are strong acids in aqueous solution better conductors of electricity than weak acids?

(vi) Polyprotic Acids

1. Using phosphoric acid as an example, define “polyprotic”.

2. Give two more examples of a polyprotic acids.

3. Write balanced equations including states to show each of the acids you mentioned in part (2) ionising in water to produce an acidic solution.
### Ionic Product of Water

\[
[H_3O^+] \times [OH^-] = 10^{-14} \text{ M}^2 \text{ at } 25^\circ C
\]

### pH Formula

\[ \text{pH} = -\log_{10}[H_3O^+] \]

### H₃O⁺ Formula

\[ [H_3O^+] = 10^{-\text{pH}} \]

### [OH⁻] Formula

\[ [H_3O^+] = \frac{10^{-14}}{[OH^-]} \]

### m-n-ratio method

<table>
<thead>
<tr>
<th>2KOH</th>
<th>+ H₂SO₄</th>
<th>→ K₂SO₄</th>
<th>+ 2H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image1.png" alt="Image" /></td>
<td><img src="image2.png" alt="Image" /></td>
<td><img src="image3.png" alt="Image" /></td>
<td><img src="image4.png" alt="Image" /></td>
</tr>
<tr>
<td><strong>m</strong> (mass)</td>
<td>5.500g</td>
<td>4.809g</td>
<td>8.542g</td>
</tr>
<tr>
<td><strong>n</strong> (no of moles)</td>
<td>0.09804</td>
<td>0.04902</td>
<td>0.04902</td>
</tr>
<tr>
<td><strong>ratio</strong> from the equation</td>
<td>2</td>
<td>1</td>
<td>1</td>
</tr>
</tbody>
</table>

**Take this route to the answer**

### c-v-n-ratio method

<table>
<thead>
<tr>
<th>2KOH</th>
<th>+ H₂SO₄</th>
<th>→ K₂SO₄</th>
<th>+ 2H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image5.png" alt="Image" /></td>
<td><img src="image6.png" alt="Image" /></td>
<td><img src="image7.png" alt="Image" /></td>
<td><img src="image8.png" alt="Image" /></td>
</tr>
<tr>
<td><strong>concentration</strong></td>
<td>1.10 M</td>
<td>0.50 M</td>
<td>?</td>
</tr>
<tr>
<td><strong>volume</strong></td>
<td>0.089</td>
<td>0.098</td>
<td>?</td>
</tr>
<tr>
<td><strong>n</strong> (no of moles)</td>
<td>0.098</td>
<td>0.049</td>
<td>0.049</td>
</tr>
<tr>
<td><strong>ratio</strong> from the equation</td>
<td>2</td>
<td>1</td>
<td>1</td>
</tr>
</tbody>
</table>

Calculations including mass-mass stoichiometry and concentration and volume of solutions; pH of strong acids and of strong bases.
Practice Questions for 2.1.3

(i) Ionic Product of Water

1. What is Brønsted-Lowry theory?

(ii) Using the pH, $\text{H}_3\text{O}^+$ and $[\text{OH}^-]$ formulae

1. Determine the pH of the following solutions:
   a) 1.20 M HCl solution
   b) 0.44 M HCl solution
   c) 0.01 M HCl solution

b) Determine the concentration of hydronium ions in solutions with the following pH:
   a) pH 3.3
   b) pH 7.0
   c) pH 1.0

c) Determine the pH of the following solutions:
   a) 3.50 M NaOH solution
   b) 1.5 M KOH solution
   c) 0.01 M NaOH solution

(iii) m-n-ratio (Mass-Mass Stoichiometry)

1. Illustrate the six reactions of acids in your notebooks. Add explanations and word equations to each one.

(iv) c-v-n-ratio (Volumetric Stoichiometry)

1. Using water as an example, explain the meaning of “amphiprotic”.
Redox reactions in aqueous solution including writing balanced equations for oxidation and reduction reactions, for example metal displacement reactions, corrosion of iron.
Practice Questions for 2.1.4

(i) Redox Half-Equations
1. Define the following words: oxidation, reduction, oxidant, reductant.
2. Classify the following half-equations as oxidation or reduction:
   a) \( \text{Cl}_2(g) + 2e^- \rightarrow 2\text{Cl}^-(aq) \)
   b) \( \text{Ni(s)} \rightarrow \text{Ni}^{2+}(aq) + 2e^- \)
   c) \( \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s) \)
   d) \( 3\text{Fe}^{2+}(aq) \rightarrow \text{Fe}(s) + 2\text{Fe}^{3+}(aq) \)

(ii) Overall Redox Equations
1. Identify the oxidant and the reductant in each of the following overall redox equations:
   a) \( \text{Cu(s)} + \text{Fe}^{2+}(aq) \rightarrow \text{Cu}^{2+} + \text{Fe}(s) \)
   b) \( 2\text{Ag}^+(aq) + \text{Sn}^{2+}(aq) \rightarrow 2\text{Ag}(s) + \text{Sn}^{4+} \)
   c) \( \text{Ca(s)} + 2\text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2(aq) + \text{H}_2(g) \)
2. Rewrite each of the equations in question 1 as two redox half-equations.

(iii) Oxidation Numbers
1. Determine the oxidation numbers of: \( \text{H}_2\text{CO}_3 \), \( \text{N}_2 \), \( \text{Zn(OH)}_4^{2-} \), \( \text{NO}_2^- \), \( \text{LiH} \), \( \text{Fe}_3\text{O}_4 \)

(iv) Balancing Redox Equations
1. Write balanced equations (and half-equations) for each of the following reactions:
   a) \( \text{NaBr} + \text{Cl}_2 \rightarrow \text{NaCl} + \text{Br}_2 \)
   b) \( \text{KMnO}_4 + \text{C}_2\text{H}_5\text{OH} \rightarrow \text{Mn}^{2+} + \text{C}_2\text{H}_4\text{O} + \text{K}^+ \)
   c) \( \text{Cr}_2\text{O}_7^{2-} + \text{Cl}^- \rightarrow \text{Cl}^{3-} + \text{Cl}_2 \)
   d) \( \text{MnO}_4^- + \text{SO}_3^{2-} \rightarrow \text{SO}_4^{2-} + \text{MnO}_4^{2-} \)

(v) Metal Reactivity Series
1. Using the reactivity series, write ionic equations for any reactions that would occur when:
   a) \( \text{Fe} \) is added to a solution containing \( \text{Cu}^{2+} \) ions
   b) \( \text{Sn} \) is added to a solution containing \( \text{Cu}^{2+} \) ions
   c) \( \text{Fe} \) is added to a solution containing \( \text{Zn}^{2+} \) ions
   d) \( \text{Mg} \) is added to a solution containing \( \text{Zn}^{2+} \) ions

(vi) Galvanic Cell
1. Define the following terms: electrode, electrolyte, anode, cathode, salt bridge, half-cell
2. Draw the following galvanic cell:
   • an \( \text{A}^{2+}(aq)/\text{A(s)} \) half-cell and a \( \text{B}^{2+}(aq)/\text{B(s)} \) half-cell are connected
   • the mass of the \( \text{A} \) electrode increases, while the mass of the \( \text{B} \) electrode decreases over time

(vii) Electrochemical Series
1. Give 5 possible pairs of electrodes for a galvanic cell. Identify the anode and the cathode.

(viii) Corrosion
CFCs

CFCs are stable, non-toxic, non-flammable compounds containing chlorine, fluorine and carbon.

CFCs were used as solvents and refrigerants until it was discovered that they break down the ozone layer in the presence of UV light.

Supercritical CO$_2$ (abbreviated scCO$_2$)

scCO$_2$ is CO$_2$ at very high pressure (73 atm). It is a safe solvent that doesn’t damage the ozone layer.

Excellent solvent used for: decaffeinating coffee, dry cleaning.

Atom Economy (we try to maximise it)

Atom Economy = \( \frac{\text{molar mass of useful product}}{\text{molar mass of all reactants}} \)

application of the principles of green chemistry; for example, replacement of halogenated solvents with supercritical carbon dioxide in industrial processes or in plant crop protection.
Practice Questions for 2.1.5

(i) Green Chemistry
Ozone Layer

In the ozone layer:

UV light breaks down \( O_3 \) and \( O_2 \)

\[
O_2 \xrightarrow{\text{UV radiation}} O + O
\]

\[
O_3(g) \xrightarrow{\text{UV radiation}} O_2(g) + O(g)
\]

Heat is released when they re-form

\[
O + O_2 \xrightarrow{} O_3
\]

\[
3O_2 \xrightarrow{} 2O_3
\]

This constant process of break-up and re-forming absorbs UV light and heats up the stratosphere

Carbon-Oxygen Cycle

- \( \text{release } CO_2 \)
- \( \text{burning, respiration, decomposition} \)
- \( \text{HUMAN ACTIVITY INCREASES THIS} \)
- \( \text{captures } CO_2 \)
- \( \text{photosynthesis} \)

Nitrogen Cycle

- \( \text{release } N \)
- \( \text{soil bacteria} \)
- \( \text{capture } N \)
- \( \text{lightning, volcanoes, nitrogen-fixing bacteria} \)
- \( \text{HUMAN ACTIVITY INCREASES THIS} \)
- \( \text{Haber process} \)

role of the atmosphere in maintaining life in the environment: effects of human activities, such as agriculture, industry, transport, energy production, on the atmosphere
Practice Questions for 2.2.1a

(i) Gases in our Atmosphere
Rain dissolves CO₂

Rain water dissolves a little CO₂ as it falls, forming weak carbonic acid:

\[ \text{MOSTLY NATURAL} \quad \text{CO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{CO}_3(aq) \]

weak acid

Rain dissolves SO₂ & SO₃ (SOₓ)

\[ \text{VOLCANOES & FOSSIL FUELS} \quad \text{SO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_3(aq) \]

acidic

\[ \text{CONVERSION OF SO}_2 \text{ BY SUNLIGHT} \quad \text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq) \]

Rain dissolves NO & NO₂ (NOₓ)

\[ \text{LIGHTNING, BACTERIA, VOLCANOES & FOSSIL FUELS} \quad 2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \]

acidic

\[ 2\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}^+(aq) + \text{NO}_3^-(aq) + \text{HNO}_2(aq) \]

nitric acid nitrous acid

Acid rain dissolves calcium carbonate (e.g. rocks and rock statues)

role of the atmosphere in maintaining life in the environment: chemical reactions and processes of acid rain
Practice Questions for 2.2.1b

(i) CO₂
(ii) SOₓ
(iii) NOₓ
Depletion of the Ozone Layer

**nitrogen(II) oxide depletes ozone**

\[
\begin{align*}
\text{NO(g)} + \text{O}_3(g) & \rightarrow \text{NO}_2(g) + \text{O}_2(g) \\
\text{NO}_2(g) + \text{O}(g) & \rightarrow \text{NO}(g) + \text{O}_2(g)
\end{align*}
\]

**CFCs deplete ozone**

- \(\text{UV radiation}\)
  - \(\text{CCl}_2\text{F}_2(g) \rightarrow \text{CClF}_2(g) + \text{Cl}(g)\)
  - \(\text{Cl}(g) + \text{O}_3(g) \rightarrow \text{ClO}(g) + \text{O}_2(g)\)
  - \(\text{ClO}(g) + \text{O}(g) \rightarrow \text{O}_2(g) + \text{Cl}(g)\)

**Net reaction**: \(\text{O}_3(g) + \text{O}(g) \rightarrow 2\text{O}_2(g)\)

An ozone hole decreases UV light absorption and causes an increase in skin cancers. International protocols have decreased CFC emissions in recent years and the ozone hole has stopped expanding.

Photochemical Smog

**CO**
- From the incomplete combustion of hydrocarbons such as fossil fuels
- CO competes with CO\(_2\) when binding to haemoglobin and can starve the body of oxygen. Can cause headaches, dizziness and death.

**SO\(_x\)**
- From burning hydrocarbons that contain traces of sulphur
- \(\text{SO}_x\) cause acid rain

**NO\(_x\)**
- From the high temperature reaction between nitrogen and oxygen in the combustion chamber of the engine
- \(\text{NO}_x\) cause acid rain
- \(\text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g)\)
- \(2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)\)

role of the atmosphere in maintaining life in the environment: qualitative effects of ozone depletion and photochemical smog
Practice Questions for 2.2.1c

(i) Ozone Depletion

(ii) Photochemical Smog
role of the atmosphere in maintaining life in the environment: role of the carbon and nitrogen cycles in maintaining life on earth
Practice Questions for 2.2.1d

(i) Carbon-Oxygen Cycle

(ii) Nitrogen Cycle
Enhanced Greenhouse Effect

The Earth has a natural greenhouse effect.

Many gases (H₂O, CO₂, CH₄, O₃, CFCs and more) all cause warming in the atmosphere because they absorb UV radiation and release it as heat. This warming is called the greenhouse effect.

The extra warming caused by gases released as a result of human activity is called the enhanced greenhouse effect.

1997 Kyoto Protocol

In 1997, the Kyoto Protocol was negotiated as a means of reducing future global temperatures by putting limits on future CO₂ emissions.

The rationale behind the Kyoto Protocol relies on three main assumptions:

- artificial CO₂ has caused significant warming in recent decades
- global warming is undesirable
- reducing CO₂ emissions is the best way to reverse global warming
Practice Questions for 2.2.2

(i) Enhanced Greenhouse Effect

(ii) 1997 Kyoto Protocol
Kinetic molecular theory

5 big assumptions:

- random motion,
- negligible volume
- exert no forces on each other
- collide elastically with each other and the container walls
- kinetic energy which is directly proportional to the absolute temperature of the gas

Gases: take up the shape and volume of any container
Pressure: force exerted on a surface by gas particles as they collide with each other and with the walls of the container
Partial Pressure: the hypothetical pressure of that gas if it alone occupied the volume of the mixture at the same temperature

Kinetic Energy vs temperature: a small increase in temperature causes a large increase in the number of particles with kinetic energy above a certain level

kinetic energy

20°C
30°C
40°C
50°C

kinetic molecular theory and its use in explaining properties of gases
Practice Questions for 2.2.3

(i) **Kinetic Molecular Theory**

1. What are the five big assumptions of kinetic molecular theory?

2. Use kinetic molecular theory to explain why gases are:
   a) compressed easily
   b) occupy all the available space of a container
   c) are miscible (they mix with each other in all ratios)
   d) have lower densities than solids

3. Sketch a graph that depicts the distribution of particle speeds in a gas.

4. When the gas depicted in question 3 is heated, explain what happens to:
   a) the area under the graph
   b) the spread of gas velocities
   c) the velocity at which the graph peaks
Gases

Temperature conversion

\[ ^\circ K = ^\circ C + 273 \]

All temperatures must be in kelvin. Absolute zero is 0°K or –273°C, at which particles stop vibrating.

Units of Pressure

1 atm

= 760 mmHg

= 1.013\times10^5 \text{ Pa}

= 101.3 kPa

= 1.013 bar

Units of Volume

1 \text{ m}^3

= 10^3 \text{ dm}^3

= 10^3 \text{ L}

= 10^6 \text{ cm}^3

= 10^6 \text{ mL}

Avogadro’s Law

Volume is proportional to number of moles

\[ \frac{V_1}{V_2} = \frac{n_1}{n_2} \]

Molar volume (\(V_m\))

volume occupied by one mole of gas at certain temperature and pressure

Two sets of temperature and pressure are commonly used. They’re called “STP” and “SLC” and their details are all in the Chemistry data booklet on page 5.

Boyle’s Law

Pressure is inversely proportional to volume

\[ P_1V_1 = P_2V_2 \]

Combined gas equation

Combination of Boyle’s & Charles’ Laws

\[ \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \]

Charles’ Law

Volume is proportional to temperature (°K)

\[ \frac{V_1}{V_2} = \frac{T_1}{T_2} \]

Ideal Gas Equation

\[ pV = nRT \]

\[ \times \text{ constant in data book} \]

Pressure \(\times\) Volume \(\times\) no. of moles \(\times\) Temperature

Calculations including those involving gas laws, molar volume (\(V_m\)) at STP and SLC, the General Gas Equation, volume-volume and mass-volume stoichiometry.
Practice Questions for 2.2.4

(i) Avogadro’s Law
1. If 1.00 mole of a gas occupies 22.4 L under certain conditions, what is the volume of:
   a) 7.00 moles?
   b) 12.4 moles?
   c) 410 moles?
   d) $1.28 \times 10^{-3}$ moles?

(ii) Boyle’s Law
1. If a gas at 1.00 atm occupies 22.4 L in volume, what volume will the gas occupy if the gas is:
   a) compressed to 2.00 atm?
   b) compressed to 50.0 atm?
   c) compressed to 1.2 atm?
   d) compressed to 5.00 atm?

(iii) Charles’ Law
1. If a gas sample at 25°C occupies 24.5 L, what volume will the gas occupy if is is:
   1. heated to 50°C?
   2. heated to 350°C?
   3. cooled to −42°C (assume that the gas doesn’t condense)
   4. heated to 4000°C?

(iv) Molar Volume
1. What is molar volume?

(v) Combined Gas Equation
1. A cylinder of gas with volume of 5.00 L is filled to a pressure of 10.0 atm. The temperature of the cylinder is 21°C. The cylinder is then taken outside, where the temperature is 41°C, and one quarter of the gas is allowed to escape. What is new pressure inside the cylinder?
2. A balloon can hold 900 mL of air before bursting. The balloon contains 750 mL of air at 0°C. The balloon is then heated very slowly. At what temperature will it burst?

(vi) Ideal Gas Equation
1. What is an “ideal gas”?
2. Why is hydrogen a more ideal gas than water vapour?
3. How many moles of gas are present in the cylinder in part (v)?
4. Assuming that the balloon is inflated to 1.50 atm, how many moles of air are in the balloon?