States of Matter

Matter is a substance that has **mass** and occupies **space**. All living and non-living things are matter. Matter can exist as a **solid**, **liquid**, or **gas**. These three forms of matter are called the **states of matter**.

Properties of Solids, Liquids, and Gases

The three states of matter have very different properties:

Property	Solid	Liquid	Gas
Shape	Fixed	Not fixed	Not fixed
Volume	Fixed	Fixed	Not fixed
Compressibility	Cannot be compressed	Cannot be compressed	Can be compressed

Substances can exist in different states of matter under different temperature and pressure conditions. Changes in temperature and pressure can change the states of matter. For example, on freezing, water becomes ice; on boiling, water becomes steam.

Kinetic Particle Theory

The differences in the properties of the states of matter can be explained based on the kinetic particle theory. The kinetic particle theory states that all matter is made up of tiny particles that are in constant random motion. The word 'kinetic' refers to motion. Moving particles have kinetic energy, hence the name 'kinetic particle theory'.

- The kinetic particle theory:
    - Describes the states of matter
    - Explains the differences in the properties of solids, liquids, and gases
    - Explains the changes of state of matter

Solids

Fixed Shape

According to the kinetic particle theory, the particles of a solid:

- Are closely packed in an orderly manner
- Are held together by very strong forces of attraction
- Have enough kinetic energy to vibrate and rotate about their fixed positions only
- Cannot move about freely

Hence, a solid has a fixed shape.

Fixed Volume

A solid cannot be compressed since its particles are already very close to one another. Thus, a solid has a fixed volume.

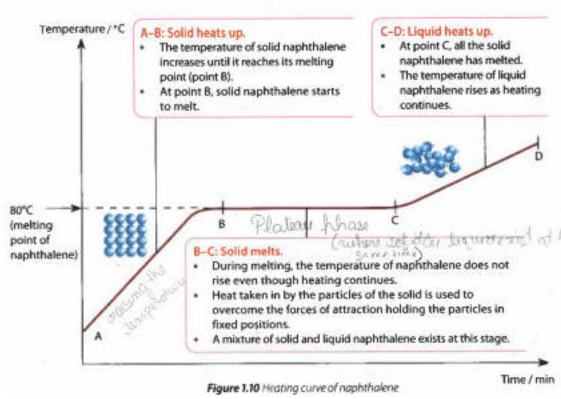


Figure 1.10 *Heating curve of naphthalene*

As seen in the diagram, the particles of a solid are closely packed together.

Liquids

No Fixed Shape

In a liquid:

- There is more space between the particles
- The particles are arranged in a disorderly manner
- The forces of attraction are weaker than in solids
- The particles have more kinetic energy than in solids
- The particles are not held in fixed positions; they move freely throughout the liquid

This is why a liquid has no fixed shape.

Fixed Volume

The particles of a liquid are further away from one another than the particles of a solid. However, the liquid particles are still packed quite closely together. Thus, a liquid cannot be compressed and has a fixed volume.

Gases

No Fixed Shape

According to the kinetic particle theory, the particles of a gas:

- Are spread far apart from one another
- Have weaker forces of attraction than the particles of a liquid
- Have more kinetic energy than the particles of a liquid
- Are not held in fixed positions; they can move about rapidly in any direction

Thus, a gas has no fixed shape.

No Fixed Volume

The particles of a gas have a lot more space between them compared to the particles of a liquid or a solid. The large space between the particles allows the gas to be easily compressed when pressure is applied. Since a gas can be compressed, it has no fixed volume.

Changes of State

Process	Change of State
Melting	Solid to liquid
Freezing	Liquid to solid
Boiling	Liquid to gas
Evaporation	Liquid to gas
Condensation	Gas to liquid

The melting point of a pure substance is the same as its freezing point.

Boiling vs. Evaporation

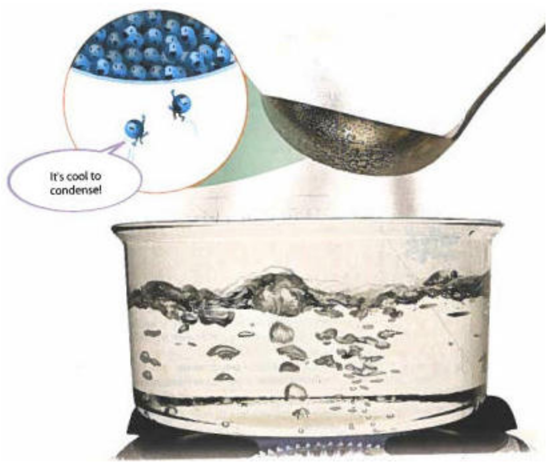
Boiling	Evaporation
Occurs only at boiling point	Occurs at temperatures below boiling point
Occurs throughout the liquid	Occurs only at the surface of the liquid
Occurs rapidly	Occurs slowly

Determining the State of a Substance

Different substances have different melting points and boiling points. If we know the melting and boiling points of a substance, we can determine whether the substance is a solid, liquid, or gas at a particular temperature.

Changes of State and Kinetic Particle Theory

Matter can change from one state to another when it is heated or cooled. Changes of state are reversible. There is no gain or loss of matter when there is a change of state.



The image shows water vapor condensing on a cold surface.

Melting

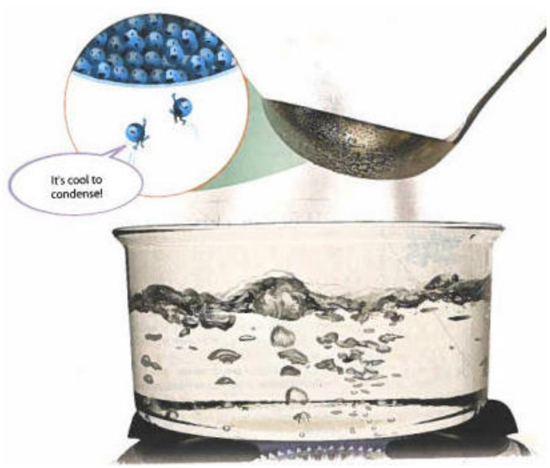


Figure 1.9 above shows what happens to the particles of a solid that is heated until it melts.

- Heat is absorbed by the particles of the solid.
- The particles start to vibrate faster about their fixed positions.
- There is an increase in their kinetic energy.
- When the temperature is high enough, the vibrations become sufficient to overcome the forces of attraction between them.
- The particles begin to break away from their fixed positions.
- The particles are no longer in their fixed positions.
- The substance is now a liquid. The particles can move freely throughout the liquid.

Boiling

Figure 1.13 shows what happens to the particles of a liquid that is heated until it boils.

- Heat is absorbed by the particles of the liquid.
- The particles start to move faster as the temperature rises.
- There is an increase in their kinetic energy.
- When the temperature is high enough, the particles have enough energy to overcome the forces of attraction holding them together.
- The particles are now spread far apart.
- The substance is now a gas. The particles can move about in any direction.

Freezing

Figure 1.11 shows what happens to the particles of a liquid that is cooled until it freezes.

- Energy is given out by the particles of the liquid.
- As the temperature is lowered, the particles no longer have enough kinetic energy to move freely.
- The particles start to settle into fixed positions.
- The particles can only vibrate about their fixed positions.
- All the particles have settled into fixed positions.
- The substance is now a solid.

Evaporation

> Evaporation occurs because some particles have enough energy to escape as a gas from the surface of a liquid. The liquid particles left behind have less kinetic energy. The average kinetic energy of the liquid particles decreases and so the average temperature of the liquid decreases.

At a higher temperature, the liquid particles have more energy. More particles have enough energy to escape as a gas from the surface of the liquid. Hence, evaporation occurs more quickly at a higher temperature. Liquids that evaporate quickly at room temperature are called **volatile liquids**. Petrol and perfumes are examples of volatile liquids.

Condensation

> During condensation, heat is given out by the gas particles. As the temperature drops, the particles lose energy and move more slowly. Eventually, the movement of the particles becomes slow enough for the gas to change to a liquid.

Effects of Temperature and Pressure on the Volume of a Gas

The particles of a gas have a lot more space between them as compared to the particles of a liquid or solid.

Effect of Temperature

When a gas is heated, its temperature increases. The particles have more energy and the space between the particles increases. The volume of the gas thus increases.

When a gas is heated, the particles have more kinetic energy. The particles collide with one another and with the wall of the container more often and at a greater force; they move further apart from one another. The higher the temperature, the faster the particles move and the greater the volume of the gas.

Effect of Pressure

When pressure is applied to a gas, the particles move closer to one another. The volume of the gas decreases.

The particles of a gas are spread far apart from one another. When we exert pressure on the gas, the particles collide with one another and with the wall of the container more often; they move closer together. Thus, when pressure is exerted on the gas, its volume decreases. If the gas particles are close enough, the forces of attraction between them become stronger and the gas will change into a liquid.

Diffusion

> Diffusion is the movement of particles from a region of higher concentration to a region of lower concentration.

When a bottle of perfume is left open for some time, the scent of the perfume soon spreads throughout the entire room. Gas particles escape from the surface of perfume. These particles move at random into the spaces between the air particles. They eventually spread throughout the entire room.

Diffusion of Liquids

Diffusion also takes place in liquids. When coffee or tea is added to water, a homogeneous mixture is formed eventually. This is because the coffee or tea particles have diffused into the spaces between the water particles.

Effect of Relative Molecular Mass on the Rate of Diffusion of Gases

The rate at which a gas diffuses depends upon its relative molecular mass.

Gas	Relative Molecular Mass	Gas	Relative Molecular Mass
Hydrogen	2	Nitrogen	28
Helium	4	Oxygen	32
Methane	16	Hydrogen chloride	36.5
Ammonia	17	Carbon dioxide	44
Carbon monoxide	28	Chlorine	71

Diffusion rates in liquids are much slower than in gases.

Worked Example

When sugar dissolves in water, the **sugar particles** move from a region of **high concentration** (the sugar lump) to a region of **low concentration** (the water). The sugar particles diffuse (slowly move) to fill up any available space between the water particles until the sugar and water particles are evenly mixed.

Factors Affecting Diffusion Rates

Gases diffuse at different rates.

Gas	Chlorine	Nitrogen	Sulfur Dioxide	Carbon Dioxide
Relative Molecular Mass	71	28	64	44

Arranging the gases according to their rates of diffusion (from slowest to fastest) yields:

Chlorine → Sulfur dioxide → Carbon dioxide → Nitrogen

Experiment: Comparing Diffusion Rates of Ammonia and Hydrogen Chloride

This experiment demonstrates the difference in diffusion rates between ammonia and hydrogen chloride gases, which have different molecular masses.

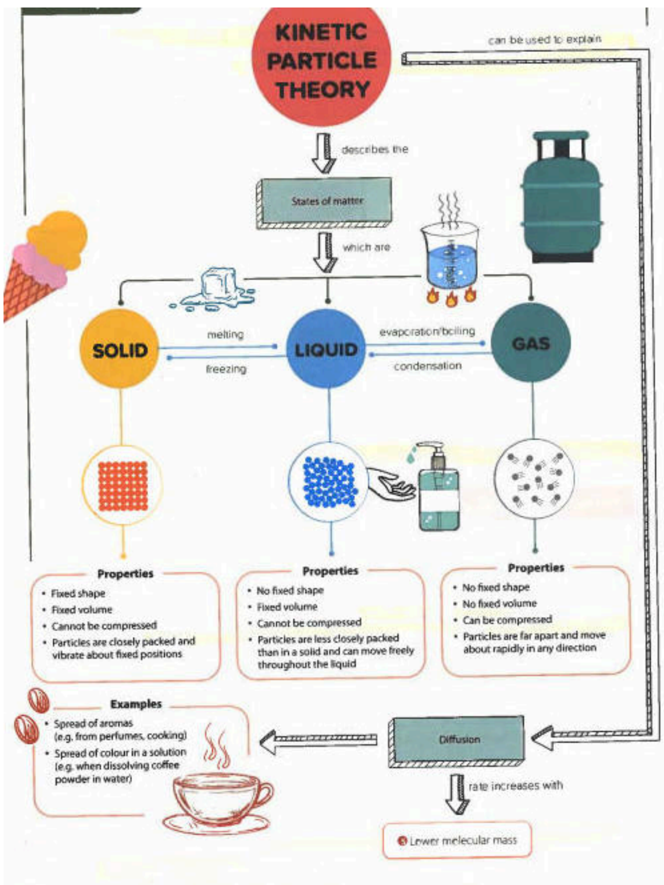


Figure 1.20 shows the experimental setup to compare the rates of diffusion of two gases. It includes cotton wool soaked in concentrated aqueous ammonia on one end of a glass tube, and cotton wool soaked in concentrated hydrochloric acid on the other end. Both ends are closed with rubber stoppers, ensuring the glass tubing is horizontal.

Procedure:

1. Soak one piece of cotton wool in concentrated aqueous ammonia and another in concentrated hydrochloric acid.
2. Place the soaked cotton wool pieces at opposite ends of a glass tube.
3. Immediately close both ends of the tube with rubber stoppers.
4. Observe where a white ring forms in the tube.

   Figure 1.21 shows the formation of a white ring due to the reaction between ammonia gas and hydrogen chloride gas. The white ring indicates the direction of diffusion of gases.

Since the white ring forms nearer to the end with hydrochloric acid, it indicates that ammonia particles move faster than hydrogen chloride particles. Ammonia gas diffuses faster than hydrogen chloride gas because ammonia has a lower molecular mass. Gases with lower molecular masses diffuse faster than those with higher molecular masses.

Quick Check

The greater the molecular mass of a gas, the slower its rate of diffusion. (True)

States of Matter Review

Multiple-Choice Questions

1. Figure 1.23 shows the particles of a substance in three states, R, S, and T.

Which of the following statements is true?

- D. The substance has a fixed volume in state S.
2. Hydraulic brakes in cars are filled with liquids and not gases. This is because gases are easily compressed, but liquids cannot be compressed. Which statement supports this explanation?
- C. The gas particles are spaced further apart than the liquid particles.
3. Condensation occurs when,
- D. A vapor turns into a liquid
4. When bubbles of gas form in a liquid, which physical change is taking place?
- C. Evaporating
5. A boiling tube containing a colorless liquid W was placed in a beaker of boiling water. Liquid W started to boil. The boiling point of W is
- C. between room temperature and 100°C

6. Jason, Siti, and Megan were discussing the kinetic particle theory.

   - Jason said that in a solid, the particles are close together.
   - Megan said that the particles of a substance in different states move at a constant speed.
   - Siti said that the higher the temperature, the faster the particles move.

Who are correct?

- B. Jason and Siti only

7. Which two statements are correct?

   - 3 The volume of a gas decreases when pressure is exerted
   - 2 The volume of a gas increases when temperature increases.
- C 2 and 3
8. A teacher was demonstrating diffusion using nitrogen dioxide, which is a brown gas that is denser than air. She inverted a gas jar of nitrogen dioxide on top of a gas jar containing air. Which option correctly describes the colors inside the gas jars after a long period of time?
- A Brown | Brown
9. Table 1.7 shows the relative molecular mass of four gases.

Gas	Carbon dioxide	Carbon monoxide	Methane	Oxygen
Relative molecular mass	44	28	16	32

Which gas will diffuse the fastest?

- C Methane

Short-Answer and Structured Questions

1. State the process in which the change of state is opposite to that in
    - (a) boiling: Condensation
    - (b) freezing: Melting
2. Figure 1.24 represents particles of a substance at room temperature. What is the state of the substance at room temperature? Explain your answer.
    - Solid. The particles are closely packed and arranged in a regular pattern, indicating a solid state.
3. Table 1.8 shows the melting and boiling points of three substances.

Substance	Sodium	Chlorine	Sodium Chloride
Melting point / °C	98	-101	801
Boiling point / °C	883	-34	1465

    - (a) State the temperature at which sodium will change from a solid to a liquid.
        - 98°C
    - (b) Deduce the state of each substance at 100°C.
        - Sodium: Liquid
        - Chlorine: Gas
        - Sodium Chloride: Solid
4. Zara left a glass of water in her room. Five days later, all the water disappeared.
    - (a) Describe what happened to the water in the glass.
        - The water evaporated, changing from a liquid to a gaseous state.
    - (b) On another occasion, Zara left the same volume of water in the same glass in her room. The water disappeared after three days. Explain why the water disappeared more quickly on this occasion.
        - The rate of evaporation was faster, possibly due to higher temperature, lower humidity, or increased air flow, which promoted quicker evaporation.
5. The melting and boiling points of five elements, A to E, are shown in Table 1.9.

Element	Melting point / °C	Boiling point / °C
A	-219	-186
B	-189	-183
C	-7	58
D	29	222
E	666	2450

    - (a) At room temperature (30°C), state which element(s) exist(s) as:
        - (i) a solid: D, E
        - (ii) a liquid: None
        - (iii) a gas: A, B
    - (b) Describe what will happen to the particles of element C when it is cooled from 80°C to -10°C.
        - As element C cools from 80°C to 58°C, it will exist as a gas. At 58°C, it will condense into a liquid. As it continues to cool to -7°C, it will remain a liquid. Finally, at -7°C, it will freeze into a solid. Throughout this process, the particles will lose kinetic energy, move slower, and become more ordered.
6. A liquid, X, was allowed to cool in air. The temperature was measured every five seconds. Figure 1.25 represents the cooling curve of X.

- (a) State the melting point of X.
  - 40°C
- (b) What is the room temperature? Explain your answer.
  - 20°C. The graph shows that the temperature of X stabilized at 20°C, indicating that the liquid has reached thermal equilibrium with its surroundings, which is the room temperature.
- (c) State the parts of the graph where energy is being given out to the surroundings.
  - All parts of the graph (A, B, C, and D) show that energy is being given out to the surroundings.
- (d) X has a boiling point of 98°C. Explain, in terms of the kinetic particle theory, what happens to the particles of X as it is heated from 10°C to 50°C.
  - As X is heated from 10°C to 40°C, the particles gain kinetic energy and move faster. At 40°C, X begins to melt. From 40°C to 50°C, the particles continue to gain kinetic energy, moving more freely as X transitions from a solid to a liquid.
- (e) Sketch a graph of temperature against time for X when it is heated from 30°C to 140°C.

7. When a person wearing perfume enters a room, the fragrance is soon smelt in other parts of the room. Explain why.

   - The perfume molecules evaporate and mix with the air molecules. These molecules then diffuse from an area of high concentration (near the person) to areas of lower concentration (other parts of the room) due to their random motion, eventually spreading the fragrance throughout the room.

Elements and Compounds

Hydrogen in Airships

Hydrogen, the lightest element on Earth, was used in airships to provide buoyancy. However, due to its flammability, disasters like the Hindenburg led to the end of airship travel. Now, scientists propose using airships for delivering goods to reduce greenhouse gas emissions.

What are Elements?

> An **element** is a pure substance that cannot be broken down into two or more simpler substances by chemical processes.

For example, carbon cannot be broken down further, making it an element. Sugar, however, can be broken down into carbon and water vapor, thus it is not an element.

`![When sugar is heated, it breaks down into carbon and water vapor](Figure: 2.1 When sugar is heated, it breaks down into carbon and water vapour)`

> Figure 2.1 shows that when sugar is heated, it breaks down into carbon and water vapor. Carbon and water vapor can then be broken down into sugar.

Chemical Symbols of Elements

Chemists use chemical symbols to represent elements. A chemical symbol consists of one or two letters. If the symbol contains two letters, the first letter is in uppercase, and the second letter is in lowercase.

Element	Symbol	Element	Symbol
Calcium	Ca	Mercury	Hg
Carbon	C	Neon	Ne
Chlorine	Cl	Oxygen	O
Hydrogen	H	Silicon	Si
Iron	Fe	Sulfur	S

Atoms

> **Atoms** are the smallest particles of an element that have the chemical properties of that element.

Each element contains only one type of atom. Noble gases like helium, neon, argon, krypton, xenon, and radon exist as individual atoms and are called **monatomic elements**.

`![Monatomic elements](Figure: 2.3 Monatomic elements)`

> Figure 2.3 shows examples of monatomic elements, including Helium (He), Neon (Ne), Argon (Ar), Krypton (Kr), Xenon (Xe), and Radon (Rn). These elements exist as individual atoms.

Molecules

> A **molecule** is a group of two or more atoms that are chemically combined.

Diatomic molecules are formed by the combination of two atoms.

`![Diatomic molecules](Figure. 24 Diatomic molecules)`

> Figure 2.4 shows diatomic molecules, which are formed by the combination of two atoms. Examples include Hydrogen (H), Nitrogen (N), Oxygen (O), Chlorine (Cl), and Iodine (I).

Molecules can also contain three or more atoms.

`![Molecules with three or more atoms](Figure: 2.5 Molecules with three or more atoms)`

> Figure 2.5 shows molecules with three or more atoms. Examples include Ozone (O3), Phosphorus (P4), and Sulfur (S8).

Compounds

> A **compound** is a pure substance containing two or more elements that are chemically combined in a fixed ratio.

For example, water (H_2O) is a compound made up of hydrogen and oxygen. The ratio of hydrogen atoms to oxygen atoms in water is always 2:1. Changing this ratio results in a different compound, like hydrogen peroxide (H_2O_2).

`![The ratio of hydrogen to oxygen in water is always 2:1](Figure: 2.6 The ratio of hydrogen to oxygen in water is always 2:1)`

> Figure 2.6 shows that water is made up of two atoms of hydrogen and one atom of oxygen, maintaining a fixed ratio of 2:1.

A compound may be made up of molecules or ions, which are electrically charged particles.

`![Water is made up of water molecules](Figure: 2.7(a) Water is made up of water molecules)`

> Figure 2.7(a) shows that water is made up of water molecules.

`![Sodium chloride is made up of ions](Figure: 2.7(b) Sodium chloride is made up of ions)`

> Figure 2.7(b) shows that sodium chloride is made up of positive sodium ions and negative chloride ions.

Properties of a Compound

A compound has different properties from its constituent elements. For example, magnesium burns in oxygen to form magnesium oxide, which has different properties from magnesium and oxygen.

`![When magnesium burns in oxygen, a white solid called magnesium oxide is formed](Figure: 2.8 When magnesium burns in oxygen, a white solid called magnesium oxide is formed)`

> Figure 2.8 shows that when magnesium burns in oxygen, a white solid called magnesium oxide is formed.

Naming of Compounds

Names of compounds containing two elements end in '-ide'. Names of compounds containing oxygen end in '-ate'.

Compound	Elements Present
Sodium chloride	Sodium, chlorine
Carbon dioxide	Carbon, oxygen
Zinc sulfide	Zinc, sulfur
Magnesium nitrate	Magnesium, nitrogen, oxygen
Copper(II) sulfate	Copper, sulfur, oxygen
Calcium carbonate	Calcium, carbon, oxygen

Chemical Formula of a Compound

The chemical formula of a compound represents the types of atoms (elements) present and their ratio.

`![Chemical formula of water](Figure: 2.9 Chemical formula of water)`

> Figure 2.9 illustrates the chemical formula of water (H2O). The symbols indicate hydrogen and oxygen atoms, and the subscript '2' shows there are two hydrogen atoms for every oxygen atom.

Compound	Chemical Formula	Ratio of Atoms
Hydrogen chloride	HCl	H:Cl = 1:1
Carbon dioxide	CO_2	C:O = 1:2
Carbon monoxide	CO	C:O = 1:1
Sulfuric acid	H_2SO_4	H:S:O = 2:1:4
Ethanol	C_2H_5OH	C:H:O = 2:6:1

For formulae containing brackets, multiply the number of atoms of each element by the subscript outside the brackets.

`![The ratio of atoms in lead(II) nitrate is lead: nitrogen: oxygen = 1:2:6](Figure: 2.10 The ratio of atoms in lead(II) nitrate is lead: nitrogen: oxygen = 1:2:6)`

> Figure 2.10 illustrates that in lead(II) nitrate, Pb(NO3)2, there is 1 lead atom, 2 nitrogen atoms (1 x 2), and 6 oxygen atoms (3 x 2). The ratio of atoms is lead: nitrogen: oxygen = 1:2:6.

Decomposition of Compounds

Chemical processes like thermal decomposition and electrolysis can break down compounds into elements or simpler compounds.

`![Thermal decomposition of copper(II) carbonate](Figure: 2.11 Thermal decomposition of copper(II) carbonate)`

> Figure 2.11 shows the thermal decomposition of copper(II) carbonate, which breaks down into copper(II) oxide and carbon dioxide upon heating.

Differences Between Elements and Compounds

Feature	Element	Compound
Composition	Contains only one type of atom	Contains two or more elements chemically combined in a fixed ratio
Smallest Particle	Atom	Molecule or ion
Decomposition	Cannot be broken down into simpler substances	Can be broken down into elements or simpler compounds by thermal decomposition or electrolysis

Let's Practice

1. Figure 2.12 shows the composition of elements in the Earth's crust.
    - (a) State the chemical symbols for oxygen and silicon.
        - Oxygen: O
        - Silicon: Si
    - (b) State the chemical names for the other elements in Figure 2.12.
        - Magnesium: Mg
        - Calcium: Ca
        - Iron: Fe
        - Aluminum: Al

!\[The composition of elements in the Earth's crust](Figure: 2.12 The composition of elements in the Earth's crust)

> Figure 2.12 illustrates the composition of elements in the Earth's crust, including oxygen, silicon, magnesium, calcium, iron, and aluminum.

2. The chemical formulae of several substances are listed: Ag, Co, CO, $CaSO_4$, Cu, KNO_3, Mn
    - (a) Name the elements present in each of the substances.
        - Ag: Silver
        - Co: Cobalt
        - CO: Carbon, Oxygen
        - $CaSO_4$: Calcium, Sulfur, Oxygen
        - Cu: Copper
        - KNO_3: Potassium, Nitrogen, Oxygen
        - Mn: Manganese
    - (b) Classify the substances as either elements or compounds.
        - Elements: Ag, Co, Cu, Mn
        - Compounds: CO, $CaSO_4$, KNO_3

3. What is the ratio of the elements in each of the following compounds?
    - To be determined by exercise 2A

Mixtures

> A **mixture** is made up of two or more substances that are not chemically combined.

Mixtures can be made up of elements or compounds. The components of a mixture are not fixed and can be present in any ratio.

!\[Different types of mixtures](Figure. 2.13 Different types of mixtures)

> Figure 2.13 shows different types of mixtures, including mixtures of elements (neon and hydrogen), mixtures of compounds (water and carbon dioxide), and mixtures of an element and a compound (hydrogen and ammonia).

Alloys are mixtures of metals with other elements (metals or non-metals).

!\[Most wind instruments such as the trumpet are made of brass which is an alloy of copper and zinc](Figure: 2.14 Most wind instruments such as the trumpet are made of brass which is an alloy of copper and zinc)

> Figure 2.14 shows that most wind instruments, such as the trumpet, are made of brass, which is an alloy of copper and zinc.

Differences Between Mixtures and Compounds

Feature	Mixture	Compound
Separation	Components can be separated by physical processes (e.g., magnetic separation, filtration)	Can only be broken down into its elements or simpler compounds by chemical processes (e.g., thermal decomposition, electrolysis)
Properties	Chemical properties are the same as those of its components	Physical and chemical properties are different from those of its constituent elements
Energy Change	Little or no energy change during formation	Usually, there is an energy change during formation
Fixed Proportion	Components can be mixed in any proportion	Elements are always combined in a fixed proportion

Determining Elements, Compounds, and Mixtures

An experiment can be conducted to determine if a substance is an element, a compound, or a mixture.

Procedure:

1. Observe the appearance of the substance.
2. Place some of the substance in a test tube of water and observe what happens.
3. Heat some of the substance strongly and observe what happens.

Substances A to E undergo the procedure, and the observations are recorded.

Substance	Appearance	Observation when water is added	Observation when heated
A	Black solid	Floats on water	Burns to give a colorless gas as the product
B	White solid		

Elements, Compounds, and Mixtures

Identifying Substances Based on Their Behavior

Let's consider how substances react when heated or mixed with water to determine whether they are **elements, compounds, or mixtures**.

Substance	Initial State	Reaction Upon Heating	Reaction with Water	Conclusion
A	No reaction occurs	Decomposes into a white solid and a gas	N/A	Element
B	Grey-green solid	Separates into grey solid that sinks in water and a yellow solid that floats on water	N/A	Compound
C	Colourless liquid	No reaction occurs	Liquid evaporates and a white solid is left behind	N/A
D	Green solid	No reaction occurs	Decomposes into a black solid and a gas	N/A
E	N/A	Glows red and a yellow-white solid is formed	N/A	Mixture

- **Substance A** is likely an **element** because it does not break down into simpler substances, and only one product is formed upon heating.
- **Substances B and E** are **compounds**. They break down into simpler substances when heated.
- **Substances C and D** are **mixtures**. The components in substance C separate when water is added, while the solid in substance D is separated from the liquid when heated.

Example: Water vs. a Hydrogen and Oxygen Mixture

Water (H_2O) is formed when hydrogen gas (H_2) and oxygen gas (O_2) are chemically combined. Let's list the differences between water and a mixture of hydrogen and oxygen:

Feature	Water (Compound)	Mixture of Hydrogen and Oxygen
Composition	Fixed ratio of hydrogen and oxygen	Variable composition
Formation	Formed with an energy change	Formed with little or no energy change
Properties	Has different properties from its elements	Has similar properties to its elements
Separation	Cannot be separated into elements by physical processes	Can be separated into its components by physical processes

Elements, Compounds, and Mixtures: Key Properties

Here is a quick recap of the general properties of elements, compounds, and mixtures:

	Elements	Compounds	Mixtures
Definition	Cannot be broken down into simpler substances by chemical processes.	Two or more elements that are chemically combined in a fixed ratio.	Two or more substances physically combined but not chemically bonded.
Formation		Are formed with an energy change.	Are formed with little or no energy change.
Composition	Have a fixed composition	Have a fixed composition	Have variable compositions.
Properties	Have general properties.	Have different properties from the elements that form it.	Have similar chemical properties as the components.
Separation	Cannot be broken down into elements or simpler compounds by chemical processes (e.g., electrolysis).	Can be separated into elements or simpler compounds by chemical processes (e.g., electrolysis).	Can be separated into its components by physical or chemical processes

Atomic Structure

Subatomic Particles

All matter is made up of **atoms**. Atoms are so tiny that they need to be magnified millions of times in order to be seen. However, atoms themselves are made up of smaller particles.

> Atoms are made up of **protons**, **neutrons**, and **electrons**. We call these smaller particles **subatomic particles**.

- A **proton** has a relative charge of +1, a relative mass of 1, and is represented by the symbol p.
- A **neutron** has a relative charge of 0, a relative mass of 1, and is represented by the symbol n.
- An **electron** has a relative charge of -1, a relative mass of 1/1840, and is represented by the symbol e.

Structure of an atom

The image above illustrates the basic structure of an atom. The protons and neutrons are tightly packed together in the **nucleus** of an atom. The electrons move rapidly around the nucleus.

Proton Number and Nucleon Number

- **Proton Number:**

> The **proton number** of an atom is the number of protons in the nucleus of an atom. The proton number is also called the **atomic number**. It is represented by the symbol Z.

Since atoms are electrically neutral, the proton number can also tell us the number of electrons in the atom. For example, the proton number of nitrogen is 7. This implies that the nitrogen atom has seven protons and seven electrons.

- **Nucleon Number:**

> The **nucleon number** of an atom is the total number of protons and neutrons in the nucleus of an atom. The nucleon number is represented by the symbol A. The nucleon number is also called the **mass number.**

$Nucleon\ number\ (A) = number\ of\ protons + number\ of\ neutrons$

Representing Proton and Nucleon Numbers

The proton and nucleon numbers can be included when representing an element in symbols:

```
 A
ZX
```

- A = nucleon number (mass number) in superscript, left of symbol
- Z = proton number (atomic number) in subscript, left of symbol
- X = chemical symbol of the element

For example, a sodium atom has 12 neutrons and 11 protons. Thus, sodium can be represented by $^{23}_{11}Na$. For convenience, it is sometimes represented using only the nucleon number, sodium-23 or ^{23}Na.

Arrangement of Electrons in Atoms

The way in which electrons are arranged around the nucleus of an atom is very important because the electron arrangement determines the chemical properties of the atom.

Electrons in an atom move around the nucleus in regions known as electron shells.

> Each electron shell corresponds to a specific energy level and can hold a certain number of electrons only.

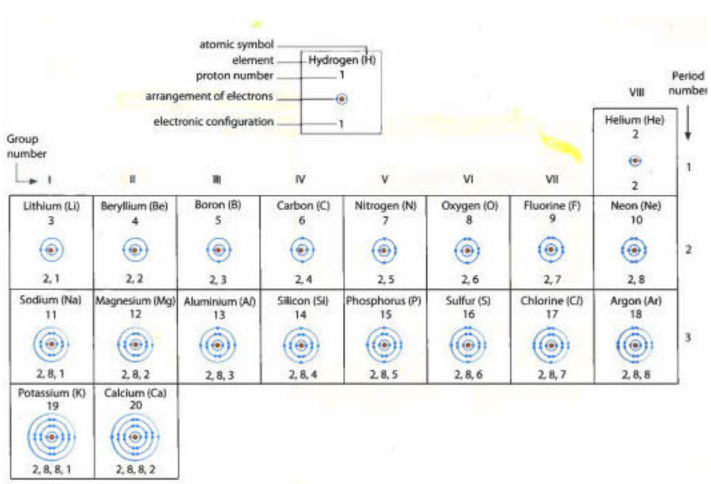
- The first shell is closest to the nucleus; corresponds to the lowest energy level; can hold a maximum of two electrons; is always filled first.
- Second and third shells have increasing energy levels (e.g., the third shell has a higher energy than the second shell); usually hold up to eight electrons each; are filled in order - usually, the second shell is fully filled before the third shell.

The arrangement of electrons in an atom can be represented using electronic configuration. For example, a magnesium atom (proton number 12) has two electrons in the first shell, eight electrons in the second shell, and two electrons in the third shell. Thus, its electronic configuration can be written as 2,8,2.

Outer Shell Electrons

> The outer shell of an atom refers to the shell that is furthest away from the nucleus of the atom. The electrons in the outer shell of an atom are known as outer electrons.

Periodic Table Connections



The image shows a section of the periodic table, highlighting the electron arrangements of elements in the first three periods. This visual aid helps to understand the relationships between electron configuration, group number, and period number.

Group VIII is often referred to as **Group 0** or the **noble gases**. The elements in this group have a full outer shell and are very unreactive.

Elements that have the same number of outer shell electrons are in the same group. The chemical properties of an element depend on the number of outer shell electrons. Thus, the elements in the same group have similar chemical properties. For example, sodium (2, 8, 1) and potassium (2, 8, 8, 1) have similar chemical properties because they each have one outer shell electron.

The number of occupied shells is equal to the period number. For example, sodium (2, 8, 1) has three occupied shells and is in period 3. Potassium (2, 8, 8, 1) has four occupied shells and is in period 4.

Isotopes

> **Isotopes** are different atoms of the same element that have the same number of protons but different numbers of neutrons.

Isotopes of the same element have the same number of protons. Thus, isotopes have the same proton number but different nucleon numbers. Most elements occur naturally as a mixture of isotopes. For example, chlorine consists of two isotopes: chlorine-35 and chlorine-37. A few elements, however, do not have isotopes. For example, all atoms of beryllium contain five neutrons and four protons.

Properties of Isotopes

Isotopes have the same chemical properties because they have the same electronic configuration. However, they have slightly different physical properties.

Property	Explanation
Chemical Properties	Chemical reactions involve only the electrons and not the neutrons. Therefore, isotopes have similar chemical properties.
Physical Properties	Physical properties are affected by mass. Isotopes have different masses and therefore different physical properties.

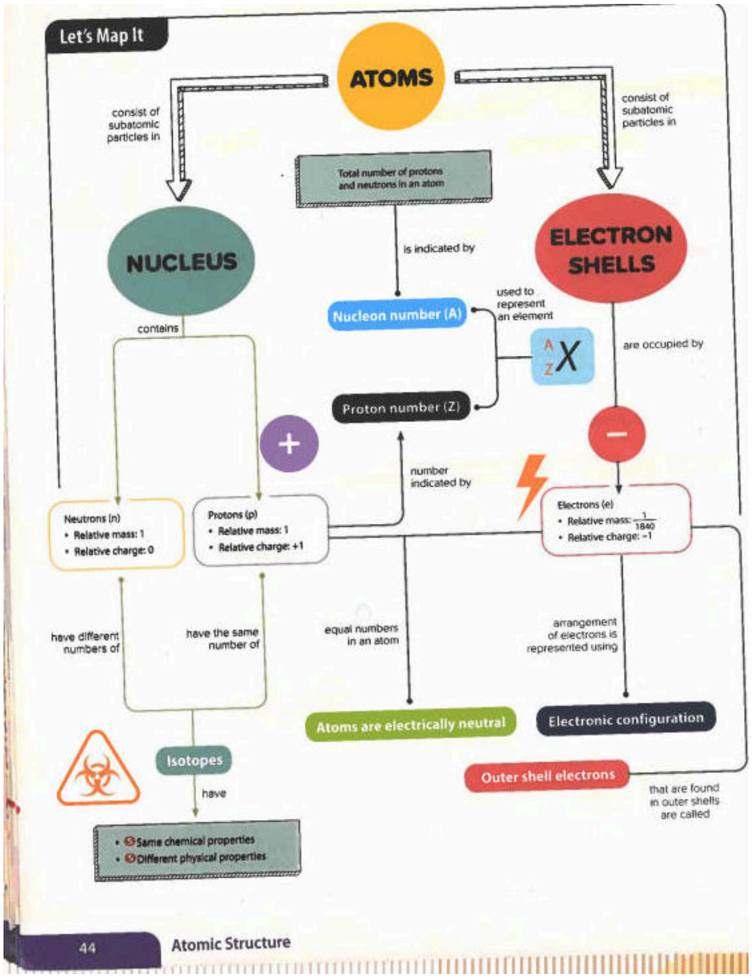
Calculating Relative Atomic Masses of an Element

The relative atomic masses of some elements are not whole numbers because these elements exist as mixtures of isotopes. The relative atomic masses of these elements are calculated based on the relative percentages of the isotopes.

For example, chlorine exists as chlorine-35 and chlorine-37. A typical sample of chlorine has 75% of chlorine-35 and 25% of chlorine-37 atoms.

Hence, the relative atomic mass of chlorine $= (\frac{75}{100} \times 35) + (\frac{25}{100} \times 37) = 26.25 + 9.25 = 35.5$.

Let's Map It



The image above is a flowchart summarizing the key concepts related to atomic structure. Starting with atoms, it branches into the nucleus and electron shells, detailing the components and properties of each. This map provides a visual overview of how atoms are composed and organized.

Isotopes, Ions, and Bonding

Isotopes

Isotopes are atoms of the same element with the same number of protons and electrons but a different number of neutrons.

Ions

Atoms are electrically neutral because they have the same number of protons and electrons. An atom becomes an ion when it gains or loses electrons, resulting in a net positive or negative charge.

- Cations: Metals form positively charged ions by losing electrons.
- Anions: Non-metals form negatively charged ions by gaining electrons.

When an atom loses electrons, it has fewer electrons than protons, forming a positive ion, or cation. The charge on a cation corresponds to the number of electrons the atom loses.

For example, a sodium atom (Na) has 11 protons and 11 electrons. When it loses one electron to achieve a noble gas electronic configuration, it forms a sodium ion (Na+) with a +1 charge, having 11 protons but only 10 electrons.

```
loses one electron
11p+ + 12n  ->  11p+ + 12n
sodium atom (Na) -> sodium ion (Na+)
2,8,1  ->  2,8
```

Magnesium (Mg) has 12 protons and 12 electrons. When a magnesium atom loses two outer electrons, it forms a magnesium ion (Mg2+).

```
loses two electrons
12p+ + 12n  -> 12p+ + 12n
magnesium atom (Mg)  ->  magnesium ion (Mg2+)
2,8,2 ->  2,8
```

Atoms of non-metals generally have more than four outer electrons. They tend to gain electrons to form anions, achieving a stable electronic configuration like a noble gas. When an atom gains electrons, it has more electrons than protons, forming a negative ion, or anion. The charge on an anion corresponds to the number of electrons the atom gains.

A chlorine atom has 17 protons and 17 electrons. To achieve a noble gas electronic configuration, the atom gains one electron to form a chloride ion.

```
gains one electron
17p+ + 18n  ->  17p+ + 18n
chlorine atom (Cl)  ->  chloride ion (Cl-)
2,8,7  -> 2,8,8
```

An oxygen atom has 8 protons and 8 electrons. To achieve a noble gas electronic configuration, the atom gains two electrons to form an oxide ion.

```
gains two electrons
8p+ + 8n ->  8p+ + 8n
oxygen atom (O)  -> oxide ion (O2-)
2,6  -> 2,8
```

Common Ions and Their Charges

Charge on cation	Name of cation	Formula of cation
+1	Sodium	Na+
+1	Potassium	K+
+1	Hydrogen	H+
+1	Ammonium	NH4+
+2	Calcium	Ca2+
+2	Copper(II)	Cu2+
+2	Magnesium	Mg2+
+2	Iron(II)	Fe2+
+3	Iron(III)	Fe3+
+3	Aluminum	Al3+
+3	Chromium(III)	Cr3+

Charge on anion	Name of anion	Formula of anion
-1	Bromide	Br^-
-1	Chloride	Cl^-
-1	Hydride	H^-
-1	Hydrogen carbonate	HCO_3^-
-1	Hydroxide	OH^-
-1	Nitrate	NO_3^-
-2	Carbonate	CO_3^{2-}
-2	Oxide	O^{2-}
-2	Sulfate	SO_4^{2-}
-3	Phosphate	PO_4^{3-}

Ionic Compounds

Ionic compounds are formed when a metal reacts with a non-metal. Metal atoms transfer their outer electrons to non-metal atoms, forming cations and anions.

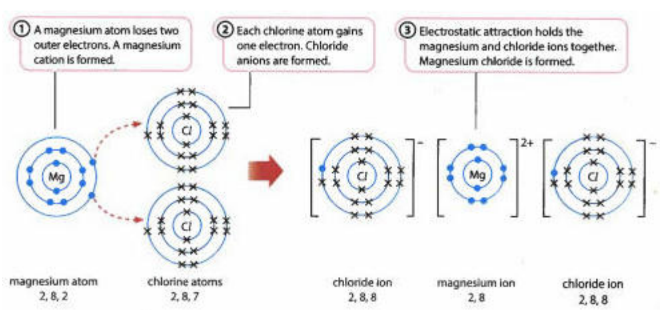
> An **ionic bond** is a strong electrostatic attraction between oppositely charged ions.

Formation of Sodium Chloride

Formation of ionic bonds in sodium chloride

The image illustrates the steps involved in forming sodium chloride (NaCl). First, a sodium atom loses its outer electron to form a sodium cation (Na^+). Simultaneously, a chlorine atom gains an electron to form a chloride anion (Cl^-). The electrostatic attraction between these oppositely charged ions holds them together, forming sodium chloride.

Formation of Magnesium Chloride



In the image, magnesium reacts with chlorine to form magnesium chloride. A magnesium atom loses two outer electrons to form a magnesium cation (Mg^{2+}), while two chlorine atoms each gain one electron to form chloride anions (Cl^-). The electrostatic attraction between the magnesium and chloride ions forms magnesium chloride ($MgCl_2$).

When magnesium reacts with oxygen, a Group VI non-metal, magnesium loses two electrons, while oxygen gains two electrons.

Dot-and-cross diagram for magnesium oxide

The image depicts the dot-and-cross diagram for magnesium oxide. Magnesium (Mg) loses two electrons to become Mg^{2+}, while oxygen (O) gains two electrons to become O^{2-}. The resulting ionic compound is held together by the electrostatic attraction between these ions.

Physical Properties of Ionic Compounds

Ionic compounds typically exhibit the following properties:

- Crystalline solids
- High melting and boiling points (not volatile)
- Good electrical conductivity when molten or in aqueous solution, but not in the solid-state

Giant Lattice Structure

Ionic compounds form giant lattice structures, which are three-dimensional networks of ions held in place by ionic bonds, packed in a regular and repeating pattern.

Giant lattice structure of sodium chloride

The strong electrostatic attractions between oppositely charged ions require a large amount of heat to overcome, resulting in high melting and boiling points.

In the solid-state, ions cannot move freely, so ionic compounds do not conduct electricity. However, when molten or dissolved in water (aqueous), the ions are free to move and conduct an electric current.

Covalent Bonding and Simple Molecules

> A **covalent bond** is formed when a pair of electrons is shared between two atoms, leading to noble gas electronic configurations.

When atoms combine by sharing electrons, molecules are formed. Covalent bonds can be formed between atoms of the same element or atoms of different elements.

Molecules of Elements

Hydrogen, chlorine, oxygen, and nitrogen exist as molecules, with each molecule made up of two atoms joined together by sharing electrons, i.e., by a covalent bond.

Formation of a covalent bond in a hydrogen molecule

In the image, two hydrogen atoms share their outer electrons to form a hydrogen molecule. Each hydrogen atom contributes one electron to the covalent bond, achieving a stable electronic configuration.

Representing a Hydrogen Molecule

Dot-and-cross diagram	**Structural formula**	**Model**
H x H or H•H	H-H	H H
Molecular formula		
H_2		

Chlorine also shares one of its outer electrons with another chlorine atom to form a covalent bond.

Representing a Chlorine Molecule

Dot-and-cross diagram	**Structural formula**	**Model**
Molecular formula		

Not all outer shell electrons may be involved in bonding. A pair of outer shell electrons not involved in bonding is called a **lone pair of electrons**.

Double and Triple Covalent Bonds

Some molecules have a double or triple covalent bond.

Formation of covalent bonds in an oxygen molecule

In the image, each oxygen atom shares two of its outer electrons with another oxygen atom, forming a double covalent bond. This arrangement allows each oxygen atom to achieve a noble gas electronic configuration.

Representing an Oxygen Molecule

Dot-and-cross diagram	Structural formula	Model	Molecular formula

The double bond in an oxygen molecule is represented by O=O in the structural formula.

![Formation of covalent bonds in a nitrogen molecule]

In the image, each nitrogen atom shares three of its outer electrons with another nitrogen atom, forming a triple covalent bond. This allows each nitrogen atom to achieve a noble gas electronic configuration.

Representing a Nitrogen Molecule

Dot-and-cross diagram	Structural formula	Model	Molecular formula

The triple bond in a nitrogen molecule is represented by N≡N in the structural formula.

Molecules of Compounds

When atoms of different elements are joined together by covalent bonding, a covalent compound or simple molecular compound is formed.

![Formation of covalent bonds in a water molecule]

The image illustrates how a water molecule is formed through covalent bonding. An oxygen atom, with six outer electrons, shares two of its electrons with two hydrogen atoms, each having one outer electron. This sharing results in two single O-H bonds, forming a water molecule.

Representing a Water Molecule

Dot-and-cross diagram	Structural formula	Model	Molecular formula

![Formation of covalent bonds in a methane molecule]

Chemical Bonds & Molecular Structures

Covalent Bond Formation

Covalent bonds are formed through the **sharing of electrons** between atoms to achieve a stable noble gas electron configuration.

Methane (CH_4)

A **methane molecule** is formed when one **carbon atom** shares its four outer electrons with four **hydrogen atoms**. Methane can be represented in different ways:

Representation	Description
Dot-and-cross diagram	Shows the sharing of electrons between carbon and hydrogen atoms.
Structural formula	Shows the arrangement of atoms and bonds in the molecule (e.g., H-C-H).
Molecular formula	Indicates the number and type of atoms present in the molecule (CH_4).
Model formula	A 3D representation showing the spatial arrangement of atoms.

Ammonia (NH_3)

An **ammonia molecule** is formed when one **nitrogen atom** shares its three outer electrons with three **hydrogen atoms**.

Representation	Description
Dot-and-cross diagram	Shows the sharing of electrons between nitrogen and hydrogen atoms.
Structural formula	Shows the arrangement of atoms and bonds (e.g., H-N-H).
Molecular formula	Indicates the number and type of atoms present in the molecule (NH_3).

Hydrogen Chloride (HCl)

A **hydrogen chloride molecule** is formed when a **hydrogen atom** shares its outer electron with a **chlorine atom.**

Representation	Description
Dot-and-cross diagram	Shows the sharing of electrons between hydrogen and chlorine atoms.
Structural formula	Shows the arrangement of atoms and bonds (H-Cl).
Molecular formula	Indicates the number and type of atoms present in the molecule (HCl).

Methanol (CH_3OH)

A **methanol molecule** is made up of **carbon, hydrogen, and oxygen atoms**.

Representation	Description
Dot-and-cross diagram	Shows the sharing of electrons between carbon, hydrogen, and oxygen atoms.
Structural formula	Shows the arrangement of atoms and bonds (e.g., H-C-O-H).
Molecular formula	Indicates the number and type of atoms present in the molecule (CH_3OH).

Carbon Dioxide (CO_2)

A **carbon atom** shares its four outer electrons with two **oxygen atoms**, forming two double covalent bonds.

Representation	Description
Dot-and-cross diagram	Shows the sharing of electrons between carbon and oxygen atoms, forming double bonds.
Structural formula	Shows the arrangement of atoms and double bonds (O=C=O).
Molecular formula	Indicates the number and type of atoms present in the molecule (CO_2).
Model formula	A 3D representation showing the spatial arrangement of atoms.

Ethene (C_2H_4)

An **ethene molecule** is a covalent compound with a double covalent bond.

Representation	Description
Dot-and-cross diagram	Shows the sharing of electrons between carbon and hydrogen atoms, with a double bond between the carbons.
Structural formula	Shows the arrangement of atoms and bonds, including the double bond (C=C).
Molecular formula	Indicates the number and type of atoms present in the molecule (C_2H_4).

Physical Properties of Simple Molecules

Simple molecules have simple molecular structures and exhibit the following properties:

- **Low melting and boiling points:**
  - Bromine (Br_2) has a melting point of -7°C and a boiling point of 59°C.
  - Atoms within each bromine molecule are held together by strong covalent bonds.
  - Weak intermolecular forces exist between the bromine molecules.
  - Not much heat is needed to break the weak intermolecular forces.
- **High volatility**
- **Poor electrical conductivity:**
  - Simple molecules do not conduct electricity in solid, liquid, or gaseous states.
  - They do not have free-moving ions or electrons to conduct electricity.

  It's a common mistake to think that covalent bonds are weak because covalent substances have low melting and boiling points. Low melting and boiling points result from weak intermolecular forces. Covalent bonds are strong and hard to break.

Giant Covalent Structures

Giant covalent structures consist of a giant network of atoms covalently bonded. They have a three-dimensional regular arrangement of atoms joined by strong covalent bonds. Examples include diamond, graphite, and silicon(IV) oxide. Diamond and graphite are **allotropes of carbon**.

Physical Properties of Giant Covalent Substances

- **High melting and boiling points:**
  - A large amount of energy is required to break the strong covalent bonds.
  - Substances with giant covalent structures are solids at room temperature.
- **Do not conduct electricity** (except graphite).
  - In giant covalent substances (except graphite), all outer electrons are used to form covalent bonds.
  - There are no delocalized electrons to conduct electricity.

Diamond

Diamond is an allotrope of carbon known for its exceptional hardness and very high melting and boiling points.

- Each carbon atom is covalently bonded to four other carbon atoms, forming a three-dimensional structure.
- It is difficult to break these strong covalent bonds.
- Diamond does not conduct electricity because all outer electrons of the carbon atoms are used for bonding.
- It is used to make the cutting edges of saw blades.

Graphite

Graphite is another allotrope of carbon that is soft, slippery, fairly unreactive, and a good conductor of electricity. It is used to make dry lubricants and inert electrodes.

- Each carbon atom is covalently bonded to three other carbon atoms, forming a continuous layer of hexagons.
- It is difficult to break these strong covalent bonds.
- Layers of carbon atoms are held loosely by weak intermolecular forces, allowing them to slide over each other.
- Each carbon atom has one outer electron that is not used to form covalent bonds, creating delocalized electrons that allow graphite to conduct electricity.

Silicon(IV) Oxide (SiO_2)

Silicon(IV) oxide, commonly known as silica, is a compound made up of silicon and oxygen.

- It has a similar structure to diamond.
- Each silicon atom is bonded to four oxygen atoms, and each oxygen atom is bonded to two silicon atoms.
- It forms a three-dimensional structure with strong covalent bonds, resulting in a high melting and boiling point.
- It does not conduct electricity because all outer electrons are used for bonding.
- It is often used to make glass due to its hardness and high melting point.

Metallic Bonding

Metallic bonding holds metal atoms together in a giant metallic lattice structure.

- Metal atoms lose their outer electrons and become positively charged ions.
- Outer electrons become delocalized and move freely between the metal ions, forming a "sea" of delocalized electrons.
- The electrostatic attraction between the positive metal ions and the sea of delocalized electrons constitutes the metallic bond.

Physical Properties of Metals

- **Good electrical conductivity:**
    - Outer electrons can break away easily from the atoms and move freely within the metal lattice.
    - The greater the number of delocalized electrons, the better the electrical conductivity.
- **Malleable and ductile:**
    - Metals can be hammered into different shapes (malleable) and drawn into wires without breaking (ductile).
    - When a force is applied, the layers of metal atoms can slide over each other through the sea of delocalized electrons.
    - The metallic bonding is not disrupted.

Chemical Formulae

Molecular Formula

> A molecular formula shows the number and type of different atoms in one molecule.

For example, the molecular formula of water is H_2O, indicating that each water molecule consists of two hydrogen atoms and one oxygen atom.

Molecule	Formula	Number of Atoms
Hydrogen	H_2	2
Oxygen	O_2	2
Ozone	O_3	3

General Rules for Writing Chemical Formulae

General Rule	Example
For compounds with both metallic and non-metallic elements, the metal is written first	Calcium oxide (CaO)
The number of atoms is written as a subscript to the right of the atom's symbol	Magnesium carbonate ($MgCO_3$)
If the compound contains more than one polyatomic ion, the formula is placed within brackets	Copper(II) hydroxide ($Cu(OH)_2$)

Empirical Formula

> The empirical formula of a compound shows the simplest whole number ratio of the different atoms or ions in a compound.

Compound	Molecular Formula	Empirical Formula
Hydrogen Peroxide	H_2O_2	HO
Ethane	C_2H_6	CH_3
Propene	C_3H_6	CH_2

When a new compound is discovered, its empirical formula is determined from the mass or percentage composition of each element. The molecular formula can then be worked out from the empirical formula.

Determining the Formula of a Simple Compound from a Model

The molecular formula of a simple molecule can be deduced based on its model.

Determining the Formula of an Ionic Compound

In an ionic compound, all positive charges must equal all negative charges. The formula can be constructed by balancing the charges on the positive ions with those on the negative ions.

- Write down the ions with their charges (e.g., X^{m+}, Y^{n-}).
- Move the values *m* and *n* diagonally (but without the charges). The chemical formula is X_nY_m.

Chemical Equations

Word Equations

Chemical reactions involve reactants converting into products. A **word equation** represents these changes using words.

> Reactants → Products

- **Reactants**: Substances at the start of the reaction.
- **Products**: Substances at the end of the reaction.

For example, the reaction between hydrogen and oxygen to form water can be represented as:

```
hydrogen + oxygen → water
```

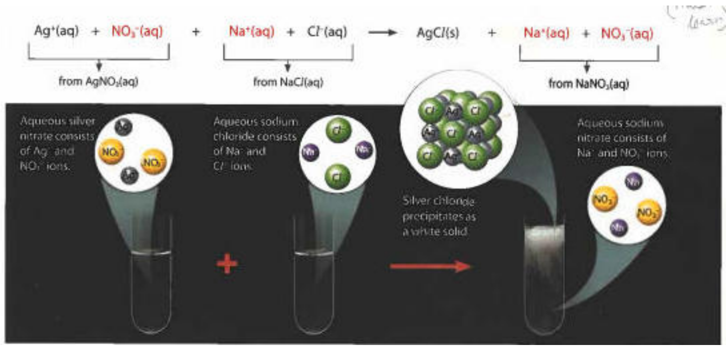
Similarly, methane reacting with oxygen to form carbon dioxide and water is:

```
methane + oxygen → carbon dioxide + water
```

Balanced Symbol Equations

A **symbol equation** or **chemical equation** uses chemical symbols to show what happens in a chemical reaction. It indicates the reactants, products, their relative amounts, and their physical states.

A **balanced symbol equation** has an equal number of atoms of each element on both sides of the equation.



The image above shows a balanced symbol equation including the reactants on the left, products on the right, state symbols indicating physical states, and numbers in front of chemical formulas to balance the equation.

Components:

- Reactants on the left-hand side.
- Products on the right-hand side.
- An arrow (→) indicating the reaction direction.
- State symbols:
  - (s) for solid
  - (g) for gas
  - (l) for liquid
  - (aq) for aqueous solution

Steps to write a balanced symbol equation:

1. Write down the chemical formulas of the reactants and products.
2. Check the number of atoms of each element on both sides.
3. Balance the equation by adding numbers in front of the chemical formulas.
4. Add the state symbols.

Ionic Equations

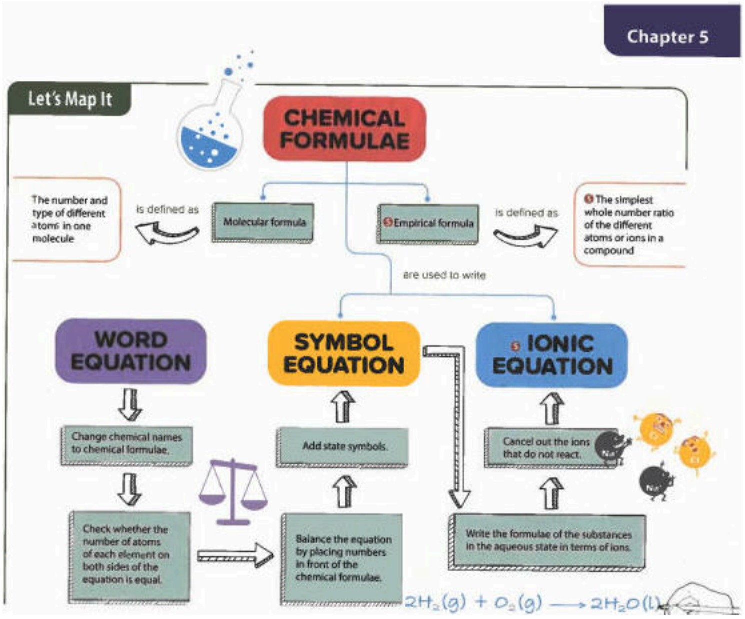
Most ionic compounds are soluble in water and exist as ions in aqueous solutions. An ionic equation is a simplified symbol equation showing reactions involving ions in aqueous solutions.

For example, when aqueous sodium chloride is added to aqueous silver nitrate, a white precipitate of silver chloride forms.

```
AgNO₃(aq) + NaCl(aq) → AgCl(s) + NaNO₃(aq)
```

In terms of ions:

```
Ag+(aq) + NO₃-(aq) + Na+(aq) + Cl-(aq) → AgCl(s) + Na+(aq) + NO₃-(aq)
```



The figure above visually represents the ionic equation, with silver and chloride ions reacting to form silver chloride precipitate.

Since only silver ions and chloride ions react, the ionic equation is:

```
Ag+(aq) + Cl-(aq) → AgCl(s)
```

Steps to write an ionic equation:

1. Write the balanced symbol equation with state symbols.
2. Rewrite the equation in terms of ions for substances in the aqueous state.
3. Cancel out the ions that do not react (spectator ions).
4. Rewrite the equation with only the ions that have reacted.

Chemical Formulae and Equations: Mind Map

Here's a summary of the key concepts related to chemical formulae and equations:

- **Word Equation:** Represents chemical reactions using words.
- **Symbol Equation:** Uses chemical symbols to show reactants, products, and their states.
- **Ionic Equation:** Shows only the ions that participate in the reaction.

The Mole

Relative Atomic Mass (Ar)

Atoms have very small masses, so scientists compare masses of different atoms relative to a standard atom. The carbon-12 (^{12}C) atom is chosen as the standard.

> The **relative atomic mass** (A_r) is the average mass of the isotopes of an element compared to $\frac{1}{12}$ of the mass of an atom of ^{12}C.

Relative atomic mass is a ratio and has no unit.

Relative Molecular Mass (Mr)

Many elements and compounds exist as molecules. The mass of a molecular substance is measured in terms of its relative molecular mass.

> The relative molecular mass (M_r) of a molecular substance is the sum of the relative atomic masses of its constituent elements.

Like relative atomic mass, relative molecular mass is a ratio and has no unit.

To calculate the relative molecular mass of a molecular substance, add together the relative atomic masses of all the atoms in its chemical formula.

Relative Formula Mass

Ionic compounds do not exist as molecules. The relative molecular mass of an ionic compound is more accurately known as the relative formula mass.

> The relative formula mass is calculated in the same way as relative molecular mass.

It is given the symbol M_r and has no unit.

The Mole (mol)

Chemists use a unit known as the mole to count particles because particles are so small and numerous.

> The mole is the unit of amount of substance. The symbol for the mole is mol.

One mole of any substance contains 6.02×10^{23} particles. The particles could be atoms, molecules, ions, or subatomic particles. The value 6.02×10^{23} is called the Avogadro constant or Avogadro number.

Calculations:

Number of moles (amount of substance) $= \frac{\text{number of particles}}{6.02 \times 10^{23}}$

Molar Mass

The molar mass of an element is the mass of one mole of atoms of the element.

- The molar mass of an element is equal to its relative atomic mass (A_r) in grams.
- The molar mass of a molecular substance is equal to its relative molecular mass (M_r) in grams.
- The molar mass of an ionic compound is equal to its relative formula mass (M_r) in grams.

Calculations involving the mole and molar mass:

Number of moles $= \frac{\text{mass of substance in grams}}{\text{molar mass of substance in g/mol}}$

Molar Mass Examples

Molar Mass of Lead

To find the mass of lead, we'll use the following steps:

1. Multiply the number of moles of lead by the molar mass of lead.

```
0.02 x 10 = number of moles of lead
```

2. Multiply the result by the molar mass of lead (`$207 \text{ g/mol}$`) to get the mass of lead.

```
0.2 x 207 = 4.14 \text{ g}
```

Relative Molecular Mass of Carbon Dioxide

To find the relative molecular mass of carbon dioxide, we'll use the following steps:

1. Calculate the number of moles of carbon dioxide using the number of molecules and Avogadro's number.

```
\text{Number of moles of } CO_2 = \frac{\text{Number of molecules}}{6.02 \times 10^{23}}
```

```
\frac{1.204 \times 10^{21}}{6.02 \times 10^{23}} = 0.002 \text{ mol}
```

2. Divide the mass of carbon dioxide by the number of moles to find the relative molecular mass.

```
\text{Relative molecular mass of } CO_2 = \frac{\text{Mass in g}}{\text{Number of moles}}
```

```
\frac{0.088 \text{ g}}{0.002 \text{ mol}} = 44
```

Number of Molecules and Atoms in Ammonia

To find the number of molecules and atoms in 0.85 g of ammonia (`$NH_3$`), follow these steps:

1. Calculate the number of moles of ammonia.

```
\text{Number of moles of } NH_3 = \frac{\text{Mass of } NH_3 \text{ in g}}{\text{Molar mass of } NH_3 \text{ in g
```

```
\frac{0.85}{14 + 3 \times 1} = 0.05 \text{ mol}
```

2. Calculate the number of molecules of ammonia.

```
\text{Number of molecules of } NH_3 = \text{Number of moles of } NH_3 \times 6.02 \times 10^{23}
```

```
0.05 \times 6.02 \times 10^{23} = 3.01 \times 10^{22}
```

3. Determine the number of atoms in one molecule of ammonia (4 atoms).
4. Calculate the total number of atoms.

```
\text{Number of atoms} = 4 \times 3.01 \times 10^{22} = 1.20 \times 10^{23}
```

Number of Ions in Magnesium Oxide

To find the number of ions in 20 g of magnesium oxide (`$MgO$`), follow these steps:

1. Calculate the number of moles of magnesium oxide.

```
\text{Number of moles of } MgO = \frac{\text{Mass of } MgO \text{ in g}}{\text{Molar mass of } MgO \text{ in g/mo
```

```
\frac{20}{24 + 16} = 0.5 \text{ mol}
```

2. Recognize that 1 mole of MgO contains 1 mole of Mg^{2+} ions and 1 mole of O^{2-} ions.
3. Determine the number of moles of ions (1 mole of ions).
4. Calculate the number of ions.

```
\text{Number of ions} = \text{Number of moles of ions} \times 6.02 \times 10^{23}
```

```
1 \times 6.02 \times 10^{23} = 6.02 \times 10^{23}
```

Molar Volume of Gases

Avogadro's Law

> **Avogadro's Law**: Equal volumes of gases, under the same conditions of temperature and pressure, contain the same number of particles.

Molar Volume

> One mole of any gas occupies 24.0 dm^3 at room temperature and pressure (r.t.p.). This volume is called the **molar volume** of a gas.

At r.t.p., equal numbers of moles of gases occupy equal volumes.

Gas	Number of moles / mol	Volume at r.t.p. / dm^3
Oxygen (O_2)	1	1 x 24 = 24
	2	2 x 24 = 48
Carbon dioxide (CO_2)	1	1 x 24 = 24
	2	2 x 24 = 48

Calculating Number of Moles

The number of moles of a gas can be determined in two ways:

1. Using mass:

```
\text{Number of moles of gas} = \frac{\text{Mass of gas in g}}{\text{Molar mass of gas in g/mol}}
```

2. Using volume at r.t.p.:

```
\text{Number of moles of gas} = \frac{\text{Volume of gas in } cm^3 \text{ at r.t.p.}}{24000 \text{ } cm^3}
```

Volume of Gas

The formula for the volume of gas can be rearranged to give:

- Volume of gas in `$cm^3$` = number of moles x `$24,000 \text{ } cm^3$`
- Volume of gas in `$dm^3$` = number of moles x `$24 \text{ } dm^3$`

Volume of Oxygen Gas Example

To find the volume of 8g of oxygen gas (`$O_2$`) at r.t.p., follow these steps:

1. Calculate the number of moles of oxygen.

```
\text{Number of moles of oxygen} = \frac{\text{Mass of oxygen in g}}{\text{Molar mass of oxygen in g/mol}}
```

```
\frac{8}{2 \times 16} = 0.25 \text{ mol}
```

2. Calculate the volume of oxygen in `$dm^3$`.

```
\text{Volume of oxygen in } dm^3 = \text{Number of moles of oxygen} \times 24 \text{ } dm^3
```

```
0.25 \times 24 = 6 \text{ } dm^3
```

Mass of Oxygen Gas Example

To calculate the mass of oxygen gas (`$O_2$`) in a room that measures 4 m high, 8 m wide, and 10 m long, assuming air contains 21% oxygen, follow these steps:

1. Calculate the volume of air in the room.

```
\text{Volume of air} = 4 \times 8 \times 10 = 320 \text{ } m^3
```

2. Convert the volume to `$cm^3$`.

```
320 \text{ } m^3 = 320 \times 10^6 \text{ } cm^3
```

3. Calculate the volume of oxygen in the air.

```
\text{Volume of oxygen} = 320 \times 10^6 \times \frac{21}{100} = 6.72 \times 10^7 \text{ } cm^3
```

4. Calculate the mass of oxygen gas.

```
\text{Mass of oxygen gas} = \frac{\text{Volume of oxygen in } cm^3}{24000 \text{ } cm^3} \times \text{Molar mass
```

```
\frac{6.72 \times 10^7}{24000} \times 32 = 8.96 \times 10^4 \text{ g}
```

Balloons and Particles

Balloons containing helium gas

The image shows balloons containing helium gas. According to Avogadro's Law, balloons with the same volume contain the same number of gaseous particles.

Balloons of the Same Mass

Equal masses of different gases do not contain the same number of particles.

Gas	Mass / g	Molar mass / g/mol	Number of moles = mass / molar mass	Number of particles = number of moles x 6.02×10^{23}
Helium (He)	0.18	4	0.18/4 = 0.045	$0.045 \times 6.02 \times 10^{23} = 2.71 \times 10^{22}$
Hydrogen (H_2)	0.18	2	0.18/2 = 0.090	$0.090 \times 6.02 \times 10^{23} = 5.42 \times 10^{22}$
Methane (CH_4)	0.18	16	0.18/16 = 0.011	$0.011 \times 6.02 \times 10^{23} = 6.62 \times 10^{21}$

Chemical Equations and Stoichiometry

Chemical Equation

A **balanced chemical equation** tells us:

- The ratio of the number of moles of the reactants and the products.
- The physical states of the reactants and products.

> **Stoichiometry**: The relationship between the number of moles of reactants and the number of moles of products involved in a chemical reaction.

Stoichiometry is based on the fact that matter cannot be destroyed or created in a chemical reaction.

Calculating Masses of Reactants and Products

Using a balanced equation, we can calculate the mass of any reactant used up or the mass of any product formed in a reaction.

Example: Sodium Hydrogen Carbonate

Calculate the mass of the solid obtained when 16.8 g of sodium hydrogen carbonate is heated strongly until there is no further change. The equation for the reaction is:

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

1. Find the number of moles of $NaHCO_3$.

```
\text{Number of moles of } NaHCO_3 = \frac{\text{Mass of } NaHCO_3 \text{ in g}}{\text{Molar mass of } NaHCO_3 \t
```

```
\frac{16.8}{23 + 1 + 12 + 3 \times 16} = 0.2 \text{ mol}
```

2. Find the mole ratio of Na_2CO_3 to $NaHCO_3$ from the equation.
   - 1 mol of Na_2CO_3 is obtained from 2 mol of $NaHCO_3$.
3. Find the number of moles of Na_2CO_3.

```
\text{Number of moles of } Na_2CO_3 = \frac{1}{2} \times \text{Number of moles of } NaHCO_3
```

```
\frac{1}{2} \times 0.2 = 0.1 \text{ mol}
```

4. Find the mass of Na_2CO_3.

```
\text{Mass of } Na_2CO_3 = \text{Number of moles} \times \text{Molar mass of } Na_2CO_3
```

```
0.1 \times (2 \times 23 + 12 + 3 \times 16) = 10.6 \text{ g}
```

Example: Iron Production

Iron is obtained from iron ore, which contains iron(II) oxide. The reaction of iron(II) oxide (Fe_3O_4) with carbon monoxide (CO) to produce iron is:

```
$Fe_3O_4(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$
```

What is the mass of iron produced when $63 \text{ } dm^3$ of carbon monoxide reacts completely with iron(II) oxide at room temperature and pressure (r.t.p.)?

1. Find the number of moles of CO.

```
\text{Number of moles of } CO = \frac{\text{Volume of } CO \text{ in } dm^3}{24 \text{ } dm^3}
```

```
\frac{63}{24} = 2.625 \text{ mol}
```

2. Find the mole ratio of Fe to CO from the equation.
   - 2 mol of Fe are produced from 3 mol of CO.
3. Find the number of moles of Fe.

```
\text{Number of moles of } Fe = \frac{2}{3} \times \text{Number of moles of } CO
```

```
\frac{2}{3} \times 2.625 = 1.75 \text{ mol}
```

4. Find the mass of Fe formed.

```
\text{Mass of } Fe = \text{Number of moles of } Fe \times \text{Molar mass of } Fe
```

```
1.75 \times 56 = 98.0 \text{ g}
```

Calculating Volumes of Gaseous Reactants and Products

For chemical reactions that involve gaseous reactants and products, the ratio of gases is the same as the volume ratio of the gases.

```
2NO(g) + O_2(g) \longrightarrow 2NO_2(g)
```

In terms of moles: 2 moles of nitrogen monoxide gas react with 1 mole of oxygen gas to form 2 moles of nitrogen dioxide gas.

In terms of volume: 2 volumes of nitrogen monoxide gas react with 1 volume of oxygen gas to form 2 volumes of nitrogen dioxide gas.

Example: Ammonia

Ammonia is manufactured by reacting nitrogen with hydrogen. Calculate the volume of nitrogen that exactly reacts with $60 \text{ } cm^3$ of hydrogen. The equation for the reaction is:

```
$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
```

1. Find the volume ratio of N_2 to H_2 from the equation.
   - 1 volume of N_2 reacts with 3 volumes of H_2.
2. Calculate the volume of N_2.

```
\text{Volume of } N_2 = \frac{1}{3} \times \text{Volume of } H_2
```

```
\frac{1}{3} \times 60 = 20 \text{ } cm^3
```

Example: Oxide of Chlorine

$10 \text{ } cm^3$ of oxygen gas reacts with chlorine gas to form $20 \text{ } cm^3$ of a gaseous oxide of chlorine. This oxide has a relative molecular mass of 87. What is the formula of this oxide?

1. Determine the volume ratio of O_2 to Cl_xO_y.
    - $1 \text{ volume of } O_2$ produces $2 \text{ volumes of } Cl_xO_y$.
2. Determine the mole ratio.
    - $1 \text{ mol of } O_2$ produces $2 \text{ mol of } Cl_xO_y$.
    - Therefore, $y = 1$.
3. Determine the formula of the oxide.

```
35.5x + 16 = 87
```

```
x = 2
```

- The formula of the oxide of chlorine is Cl_2O.

Limiting Reactants

> Limiting Reactant: The reactant that is completely used up in a reaction, limiting the amount of products formed. The reactants that are not used up in a reaction are said to be in excess.

Example

Consider the following reaction:

```
$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$
```

Experiment	Number of moles of Hydrogen (H_2)	Number of moles of Chlorine (Cl_2)	Number of moles of Hydrogen chloride (HCl)	Unreacted Hydrogen (H_2)	Unreacted Chlorine (Cl_2)
A	1	1	2	-	-
B	3	1	2	2	-
C	1	3	2	-	2

In Experiment A, the exact amounts of reactants required are used. In experiments B and C, an excess of one reactant is used.

Example: Zinc and Sulfur

```
Zn(s) + S(s) \rightarrow ZnS(s)
```

When 13.0 g of powdered zinc was heated with 8.0 g of powdered sulfur, a vigorous reaction occurred, and zinc sulfide was formed.

1. Identify the limiting reactant in this reaction.
2. Calculate the mass of the excess reactant.

(a) Find the number of moles of each reactant.

```
\text{Number of moles of zinc} = \frac{\text{Mass of zinc in g}}{\text{Molar mass of zinc in g/mol}}
```

```
\frac{13.0}{65} = 0.2 \text{ mol}
```

```
\text{Number of moles of sulfur} = \frac{\text{Mass of sulfur in g}}{\text{Molar mass of sulfur in g/mol}}
```

```
\frac{8.0}{32} = 0.25 \text{ mol}
```

(b) Find the mole ratio of the reactants from the equation to determine the limiting reactant.

- 1 mol of zinc reacts with 1 mol of sulfur.
- Therefore, 0.2 mol of zinc will react with 0.2 mol of sulfur.
- Since 0.25 mol of sulfur was used, sulfur must be in excess. The limiting reactant is zinc.
- Number of moles of unreacted sulfur = $0.25 - 0.20 = 0.05 \text{ mol}$
- Mass of unreacted sulfur = number of moles of unreacted sulfur x molar mass of sulfur
- $0.05 \times 32 = 1.6 \text{ g}$

Example: Ethene

```
$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
```

In an experiment, $10 \text{ } cm^3$ of ethene was burnt in $50 \text{ } cm^3$ of oxygen.

1. Which gas was in excess?
2. Calculate the volume of the reactant in excess at the end of the reaction.
3. Calculate the volume of carbon dioxide produced.

(a) Find the volume ratio of ethene to oxygen from the equation.

- 3 volumes of O_2 react with 1 volume of C_2H_4.

```
\text{Volume of } O_2 = 3 \times \text{Volume of } C_2H_4 = 3 \times 10 = 30 \text{ } cm^3
```

However, $50 \text{ } cm^3$ of O_2 was supplied. Oxygen gas was in excess.

(b) Volume of excess O_2 = initial volume - volume used

```
50 - 30 = 20 \text{ } cm^3
```

(c) Find the volume ratio of C_2H_4 (limiting reactant) to carbon dioxide from the equation.

- 2 volumes of CO_2 are produced from 1 volume of C_2H_4.

```
\text{Volume of } CO_2 = 2 \times \text{Volume of } C_2H_4 = 2 \times 10 = 20 \text{ } cm^3
```

The Concentration of a Solution

> Concentration: The amount of a solute dissolved in a unit volume of the solution.

Units of Concentration

Concentration can be measured in:

- grams of solute per $dm^3 \text{ (g/dm}^3)$` of solution.
- moles of solute per $dm^3 \text{ (mol/dm}^3)$` of solution (molar concentration).

Formulae

- Concentration in $g/dm^3 = \frac{\text{mass of solute in g}}{\text{volume of solution in } dm^3}$
- Mass of solute in g = concentration in $g/dm^3 \times \text{volume of solution in } dm^3$
- Concentration in $mol/dm^3 = \frac{\text{number of moles of solute}}{\text{volume of solution in } dm^3}$
- Concentration in $mol/dm^3 = \frac{\text{concentration in } g/dm^3}{\text{molar mass of solute in g/mol}}$

Examples

Paracetamol in Medicine

A bottle of medicine contains 120 mg of paracetamol in $5 \text{ } cm^3$ of solution. What is the concentration, in g/dm^3, of paracetamol in the medicine?

- $1 \text{ } dm^3 = 1000 \text{ } cm^3$
- $\text{Volume of solution} = \frac{5}{1000} = 0.005 \text{ } dm^3$

```
\text{Concentration of paracetamol in } g/dm^3 = \frac{\text{mass of paracetamol in g}}{\text{volume of solution in }
```

```
\frac{120 \times 10^{-3}}{0.005} = 24 \text{ } g/dm^3
```

Sodium Hydroxide Solution

A solution of sodium hydroxide was prepared by dissolving 3.5 g of sodium hydroxide ($NaOH$) in distilled water and making the volume up to $100 \text{ } cm^3$. What is the concentration of the sodium hydroxide solution in mol/dm^3?

1. Number of moles of $NaOH$:

   ```
   \text{mass of NaOH in g} / \text{molar mass of NaOH in g/mol}
   ```

   ```
   3.5 / (23 + 16 + 1) = 0.0875 \text{ mol}
   ```

2. Concentration of $NaOH$ in mol/dm^3:

   ```
   \text{number of moles of NaOH} / \text{volume of solution in } dm^3
   ```

   ```
   0.0875 / (100/1000) = 0.875 \text{ mol/dm}^3
   ```

Hydrochloric Acid

The concentration of a bottle of hydrochloric acid (HCl) is $73 \text{ } g/dm^3$. How many moles of HCl are there in $250 \text{ } cm^3$ of this acid?

1. Concentration of `$HCl$` in `$mol/dm^3$`:

   ```
   \text{concentration of HCl in } g/dm^3 / \text{molar mass of HCl in g/mol}
   ```

   ```
   73 / (1 + 35.5) = 2 \text{ mol/dm}^3
   ```

2. Number of moles of `$HCl$`:

   ```
   \text{volume of HCl in } dm^3 \times \text{concentration of HCl in mol/dm}^3
   ```

   ```
   (250/1000) \times 2 = 0.5 \text{ mol}
   ```

Changing the Concentration of a Solution

We can decrease the concentration of a solution by adding more solvent. We can also increase the concentration of a solution by increasing the amount of solute particles in the solution.

Volumetric Analysis

> **Volumetric analysis**: A method used to check the concentration of substances by titration.

Titration

> **Titration**: A method involving a solution of unknown concentration titrated against a solution of known concentration (titrant).

Procedure

To do volumetric analysis, we use a method called titration. A titration involves a solution of unknown concentration titrated against a solution of known concentration (titrant). In a titration, we find out the volume of titrant required to react with a fixed volume of the unknown. From these results, we can calculate the concentration of the unknown solution.

Let's Investigate 6A

Objective

To determine the concentration of ammonia in a household ammonia solution

Procedure

1. Introduce 25.0 cm^3 of ammonia of unknown concentration indicator into a conical flask using a pipette.
2. Place the sulfuric acid solution of known concentration (titrant) in a burette.
3. Add the titrants slowly into the conical flask. Swirl the flask constantly as the titrant is added.
4. Stop the titration at the end-point (when the indicator has changed color permanently, indicating that the reaction is complete). Record the volume of titrant used.

Example: Ammonia Solution Analysis

A household ammonia solution was analyzed to determine its ammonia content. `$25.0 \text{ } cm^3$` of the ammonia solution required `$21.90 \text{ } cm^3$` of `$0.11 \text{ } mol/dm^3$` sulfuric acid to achieve the end-point of titration. Calculate the concentration of ammonia, in `$g/dm^3$`, in the household ammonia solution. The equation for the reaction is:

```
$2NH_3(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$
```

1. Find the number of moles of `$H_2SO_4$`:

   ```
   \text{Volume of } H_2SO_4 (dm^3) \times \text{concentration of } H_2SO_4 mol/dm^3
   ```

   ```
   (21.90/1000) \times 0.11 = 2.409 \times 10^{-3} \text{ mol}
   ```

2. Find the number of moles of `$NH_3$` using the mole ratio from the equation:

   ```
   2NH_3(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)
   ```

   - 2 moles of `$NH_3$` react with 1 mole of `$H_2SO_4$`.

   ```
   \text{Number of moles of } NH_3 = 2 \times \text{number of moles of } H_2SO_4 = 2 \times 2.409 \times 10^{-3} = 4
   ```

3. Find the concentration of `$NH_3$` in `$mol/dm^3$`:

   ```
   \frac{\text{number of moles of } NH_3}{\text{volume of solution in } dm^3}
   ```

   ```
   \frac{4.818 \times 10^{-3}}{25.0/1000} = 0.1927 \text{ } mol/dm^3
   ```

4. Convert this to concentration in `$g/dm^3$`:

   ```
   \text{concentration in } g/dm^3 = \text{concentration in } mol/dm^3 \times \text{molar mass of } NH_3
   ```

   ```
   0.1927 \times (14 + 3) = 3.276 \text{ } g/dm^3
   ```

Mole Ratios and Calculations

To find the number of moles of NH_3 in a reaction, we can use the mole ratio from the balanced equation.

For example:

$2NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$

From the equation, 2 moles of NH_3 react with 1 mole of H_2SO_4. Therefore:

$\text{Number of moles of } NH_3 = 2 \times \text{Number of moles of } H_2SO_4$

If we have 2.409 x 10^{-3} moles of H_2SO_4, the number of moles of NH_3 is:

$2 \times 2.409 \times 10^{-3} = 4.818 \times 10^{-3}$ mol

Concentration Calculations

Concentration can be expressed in mol/dm^3.

$\text{Concentration of } NH_3 \text{ in } mol/dm^3 = \frac{\text{Number of moles of } NH_3}{\text{Volume of solution in } dm^3}$

If we have 4.818×10^{-3} moles of NH_3 in 25.0 cm^3 of solution:

$\text{Concentration of } NH_3 = \frac{4.818 \times 10^{-3}}{\frac{25.0}{1000}} = 0.1927 \text{ mol/dm}^3$

Concentration can also be expressed in g/dm^3.

$\text{Concentration of } NH_3 \text{ in } g/dm^3 = \text{Concentration of } NH_3 \text{ in } mol/dm^3 \times \text{Molar mass of } NH_3$

$\text{Concentration of } NH_3 = 0.1927 \times (14 + 3 \times 1) = 3.28 \text{ g/dm}^3$

Acid-Base Titrations

Acid-base titrations involve the use of an acid, such as hydrochloric acid (HCl), and a base, such as sodium hydroxide ($NaOH$). Indicators are used to determine when the reaction is complete. Indicators show different colors in acidic and alkaline solutions.

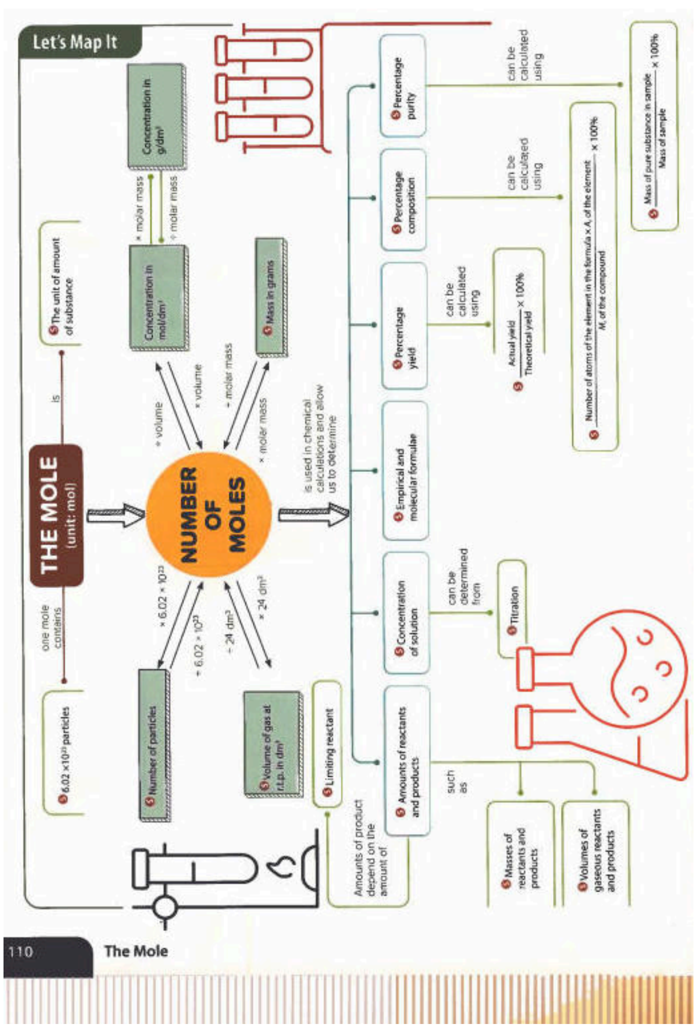


Figure 6.5: Indicators like methyl orange help determine the end-point of a titration by changing colors at different pH levels.

Example Calculation: Metal Carbonate

6. 60 g of a metal carbonate, M_2CO_3, was dissolved in 1000 cm^3 of water. 25.0 cm^3 of this solution required 30.00 cm^3 of 0.201 mol/dm^3 HCl for complete reaction. The chemical equation for this reaction is:

$$M_2CO_3(aq) + 2HCl(aq) \rightarrow 2MCl(aq) + H_2O(l) + CO_2(g)$$

(a) Calculate the concentration, in mol/dm^3, of the M_2CO_3 solution. (b) Calculate the relative atomic mass of M, and identify the metal M using the Periodic Table.

Solution

(a) First, find the number of moles of HCl used:

$$\text{Number of moles of } HCl = \frac{30.0}{1000} \times 0.201 = 6 \times 10^{-3} \text{ mol}$$

From the equation, 1 mol of M_2CO_3 reacts with 2 mol of HCl. Therefore:

$$\text{Number of moles of } M_2CO_3 \text{ in } 25.0 \text{ cm}^3 = \frac{1}{2} \times \text{Number of moles of } HCl$$

$$\text{Number of moles of } M_2CO_3 = \frac{1}{2} \times 6 \times 10^{-3} = 3 \times 10^{-3} \text{ mol}$$

$$\text{Concentration of } M_2CO_3 = \frac{3 \times 10^{-3}}{\frac{25.0}{1000}} = 0.12 \text{ mol/dm}^3$$

(b) Relative formula mass of M_2CO_3:

$$\text{Relative formula mass of } M_2CO_3 = \frac{\text{mass of } M_2CO_3}{\text{number of moles of } M_2CO_3} = \frac{16.60}{0.12} = 138.33$$

Let the relative atomic mass of M be y. Then:

$$2y + 60 = 138.33 \; 2y = 78.33 \; y = 39.2$$

Using the Periodic Table, M is potassium (K).

Empirical and Molecular Formulae

Empirical Formula

> The **empirical formula** of a compound shows the types of elements present in the compound and the simplest whole number ratio of the different atoms or ions in the compound.

To determine the empirical formula of a compound:

1. Find the mass of the reactants.
2. Work out the relative numbers of moles (mole ratio) of the reactants used.
3. Find the formula of the compound using the mole ratio obtained.

Molecular Formula

> The **molecular formula** is the actual formula of a compound.

The molecular formula is always a multiple of the empirical formula.

$$\text{If empirical formula} = A_xB_y, \text{ molecular formula} = (A_xB_y)_n \text{ where } n = 1, 2, 3, \ldots$$

To find n:

$$n = \frac{\text{relative molecular mass}}{\text{relative mass from empirical formula}}$$

Example: Glycerol

Glycerol contains 39.1% carbon and 52.2% oxygen, with the remainder being hydrogen. What is the molecular formula of glycerol? (One mole of glycerol weighs 92.0 g)

Element	C	H	O
Percentage in compound	39.1	100 - 39.1 - 52.2 = 8.7	52.2
Relative atomic mass	12	1	16
Number of moles	$\frac{39.1}{12} = 3.26$	$\frac{8.7}{1} = 8.7$	$\frac{52.2}{16} = 3.26$
Mole ratio	$\frac{3.26}{3.26} = 1$	$\frac{8.7}{3.26} = 2.67$	$\frac{3.26}{3.26} = 1$
Simplest ratio	$1 \times 3 = 3$	$2.67 \times 3 = 8$	$1 \times 3 = 3$

Empirical formula of glycerol is $C_3H_8O_3$.

Let the molecular formula of glycerol be $(C_3H_8O_3)_n$. Relative mass of glycerol from empirical formula = $(3 \times 12) + (8 \times 1) + (3 \times 16) = 92$

$$n = \frac{\text{relative molecular mass of glycerol}}{\text{relative mass of glycerol from empirical formula}} = \frac{92}{92} = 1$$

Molecular formula of glycerol = $C_3H_8O_3$

Percentage Yield, Composition, and Purity

Percentage Yield

> The **theoretical yield** of a reaction is the calculated amount of product that would be obtained if the reaction were to be complete. The **actual yield** is the amount of pure product that is actually produced in the experiment.

$$\text{Percentage Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Percentage Composition

> **Composition** refers to the substances that the compound is made up of.

$$\text{Percentage by mass of an element in a compound} = \frac{\text{number of atoms of the element in the formula} \times A_r \text{ of the element}}{M_r \text{ of the compound}} \times 100$$

Percentage Purity

> The **percentage purity** indicates how much of a substance in a sample is the desired pure compound.

$$\text{Percentage purity} = \frac{\text{mass of pure substance in sample}}{\text{mass of sample}} \times 100$$

Electrolysis and Electrochemical Reactions

Enrichment Activity: Electrical Energy to Chemical Energy

Electrical energy is converted into **chemical energy** as chemical reactions take place at the electrodes. The terms **anode** and **cathode** were introduced by scientist Michael Faraday after discovering the movement of electrons during **electrolysis**.

> The term 'anode' comes from the Latin term *anodos*, and 'cathode' comes from the Latin term *kathodos*.

- **Cations** (positive ions) gain electrons at the negatively-charged **cathode**, which is **reduction**.
- **Anions** (negative ions) lose electrons at the positively-charged **anode**, which is **oxidation**.

When cations or anions gain or lose electrons at the electrodes, they form atoms or molecules and are said to be **discharged**.

7.2 Electrolysis of Molten Ionic Compounds

- **Binary Compound**: A compound containing only two elements, usually a metal cation and a non-metal anion.
- When a molten binary compound undergoes electrolysis, a metal and a non-metal are formed as products.

Electrolysis of Molten Lead(II) Bromide

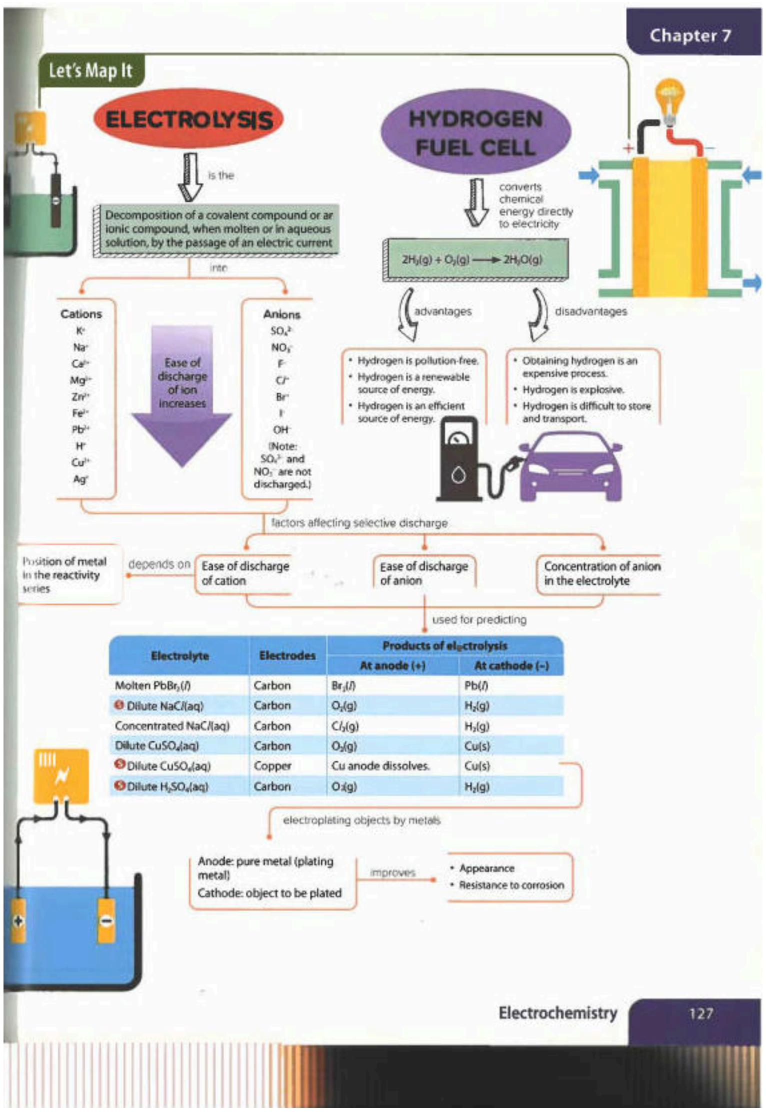


Figure 7.4 illustrates the chemical reactions occurring at the electrodes during the electrolysis of lead(II) bromide.

- **At the anode:**
  - Negatively-charged Br^- ions are attracted to the anode.
  - $2Br^- \rightarrow Br_2(g) + 2e^-$
  - Br^- ions lose electrons to form bromine gas; they are oxidized and said to be discharged.
- **At the cathode:**
  - Positively-charged Pb^{2+} ions are attracted to the cathode.
  - $Pb^{2+}(l) + 2e^- \rightarrow Pb(l)$
  - Each Pb^{2+} ion gains two electrons to form a lead atom; it is reduced and said to be discharged.

The overall reaction:

$PbBr_2(l) \rightarrow Pb(l) + Br_2(g)$

Inert Electrodes

Inert carbon electrodes are used to prevent bromine from reacting with the electrode, as bromine produced is very reactive. Inert electrodes are electrodes that do not react with the products of electrolysis or the electrolyte. Carbon (graphite) and platinum electrodes are considered inert because they are rarely involved in electrolytic reactions. Pencil lead contains graphite and can be used as inert electrodes in a simple setup for electrolysis.

7.3 Electrolysis of Aqueous Solutions of Compounds

- Metals or hydrogen gas are formed at the cathode.
- Non-metals (other than hydrogen) are formed at the anode.
- Aqueous Solution: A mixture of a compound with water.

For the electrolysis of molten compounds, it is easy to work out the reactions that occur at the electrodes because there are only two different ions present, and both will be discharged at the electrodes. However, the electrolysis of an aqueous solution is more complex. An aqueous solution of a compound is a mixture of the compound with water. For example, an aqueous solution of sodium chloride contains sodium chloride and water. Thus, the ions present in the solution are ions from sodium chloride ($Na^+(aq)$ and $Cl^-(aq)$) and the dissociation of water ($H^+(aq)$ and $OH^-(aq)$). These ions thus compete for discharge at the electrodes. In general, we can say that metals or hydrogen gas are formed at the cathode and that non-metals (apart from hydrogen) are formed at the anode.

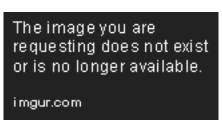
Selective Discharge of Ions

In an aqueous solution, more than one type of cation or anion is present in the electrolyte. However, only one cation and one anion will be discharged during electrolysis. This is known as selective discharge of ions. If inert electrodes are used during electrolysis, the ions discharged and hence the products formed depend on three factors:

1. Selective discharge of cations
2. Selective discharge of anions
3. Effect of concentration on the selective discharge of anions

1. Selective Discharge of Cations

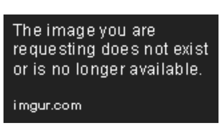
More reactive metals have a greater tendency to form ions. Thus, in electrolysis, ions of reactive metals like potassium and sodium will remain as ions and will not be discharged; ions of hydrogen and less reactive metals like copper and silver will accept electrons more readily and get discharged. The ease of discharge of cations will depend on the position of the metal (producing the cation) in the reactivity series.



Ions of metals below hydrogen in the series will be discharged during electrolysis.

2. Selective Discharge of Anions

The ease of discharge of anions:



OH^- ions give up electrons most readily during electrolysis to form water and oxygen:

$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$

3. Effect of Concentration on the Selective Discharge of Anions

An anion in higher concentration is more readily discharged. For example, in the electrolysis of concentrated aqueous sodium chloride, two different anions are attracted to the anode: Cl^- and OH^- ions. According to their relative ease of discharge, OH^- ions should be discharged preferentially. However, in concentrated aqueous sodium chloride, Cl^- ions are far more numerous than OH^- ions. Therefore, Cl^- ions are discharged at the anode in preference to OH^- ions.

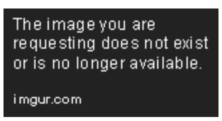
$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$

> Preferentially: In favor of something else.

Rules for Predicting Products of Electrolysis

Rule	Description
1	Identify the cations and anions in the electrolyte. Remember that an aqueous solution also contains H^+ and OH^- ions from the dissociation of water molecules.
2	Determine the anion that is discharged at the anode. The product of **electrolysis** is always oxygen unless the electrolyte contains a high concentration of Cl^-, Br^- or I^- anions.
3	Determine the cation that is discharged at the cathode. Reactive metals such as sodium and potassium are never produced during **electrolysis** of their aqueous solutions. If the cations come from a metal above hydrogen in the reactivity series, then hydrogen will be produced.
4	Identify the cations and anions that remain in the solution after **electrolysis**. They form the product remaining in the solution.

Electrolysis of Dilute Aqueous Sodium Chloride

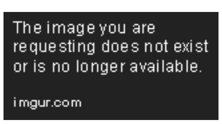


- Ions present: $Na^+(aq)$, $Cl^-(aq)$, $H^+(aq)$, and $OH^-(aq)$.
- At the anode: OH^- ions are preferentially discharged to form water and oxygen gas.
  - $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$
- At the cathode: H^+ ions are preferentially discharged as hydrogen gas.
  - $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Overall Reaction:

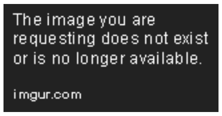
$$2H_2O(l) \xrightarrow{electrolysis} 2H_2(g) + O_2(g)$$

Electrolysis of Concentrated Aqueous Sodium Chloride



- Ions present: $Na^+(aq)$, $Cl^-(aq)$, $H^+(aq)$, and $OH^-(aq)$.
- At the anode: Cl^- ions are discharged as chlorine gas due to their higher concentration.
  - $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
- At the cathode: H^+ ions are preferentially discharged as hydrogen gas.
  - $2H^+(aq) + 2e^- \rightarrow H_2(g)$
- The remaining Na^+ and OH^- ions recombine to form sodium hydroxide, resulting in an alkaline solution.

Electrolysis of Aqueous Copper(II) Sulfate Using Inert Electrodes



- Ions present: $Cu^{2+}(aq)$, $SO_4^{2-}(aq)$, $H^+(aq)$, and $OH^-(aq)$
- At the anode: OH^- ions are preferentially discharged to give oxygen gas.
    - $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$
- At the cathode: Cu^{2+} ions are preferentially discharged as copper metal (atoms).
    - $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
- The remaining H^+ and SO_4^{2-} ions recombine to form sulfuric acid.

Electrolysis of Water and Dilute Sulfuric Acid

Pure water is a poor conductor of electricity. However, if a small amount of an ionic compound or dilute sulfuric acid is added to water, the solution becomes a good conductor of electricity.

- Dilute Sulfuric Acid: Provides hydrogen ions and sulfate ions ($H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$).
- Water: Provides hydrogen ions and hydroxide ions ($H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$).

At the cathode, the hydrogen ions are discharged, and hydrogen gas is produced:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

At the anode, the hydroxide ions are discharged in preference to the sulfate ions. Oxygen gas is produced:

$$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$$

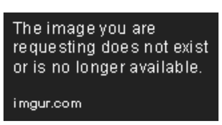
The products of the electrolysis of water are always two volumes of hydrogen at the cathode and one volume of oxygen at the anode.

7.4 Industrial Applications of Electrolysis

Electrolytic Purification of Copper

When metals are extracted from their ores, they are not pure. Thus, electrolysis is used to purify them in a process known as electrolytic purification.

Electrolysis of Aqueous Copper(II) Sulfate Using Reactive Copper Electrodes



- At the reactive copper anode:
    - $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
    - The copper anode dissolves, decreasing in mass.
- At the reactive copper cathode:
    - $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
    - The cathode becomes coated with a layer of reddish-brown copper, increasing in mass.

Electroplating

Electroplating is the process of depositing a layer of metal on another substance using electrolysis.

Uses of Electroplating

- Coating a metal object with another metal for a good decorative finish and to prevent rusting.
- Silver, chrome, or gold plating is used to coat a relatively cheap metal to make it look more expensive and to protect it from corroding.

How Objects Are Electroplated

The object to be plated is made the cathode of an electrolytic cell, and the anode is the source of the plating metal. The electrolyte is an aqueous solution of a salt of the plating metal. The net result is the transfer of the plating metal from the anode to the cathode.

7.5 Hydrogen-Oxygen Fuel Cells

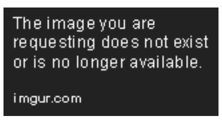
A **fuel cell** is a chemical cell in which reactants (usually a fuel and oxygen) are continuously supplied to produce electricity directly. The best-known example of a fuel cell is the **hydrogen-oxygen fuel cell**.

Hydrogen-Oxygen Fuel Cell

A hydrogen-oxygen fuel cell is a chemical cell that uses hydrogen as the fuel.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ (ΔH = -484 kJ)

This reaction is used to produce electrical energy in the hydrogen-oxygen fuel cell.

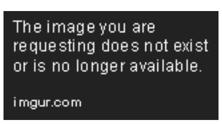


- **At the positive electrode (cathode)**: Oxygen is reduced to form hydroxide ions.
  - $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
- **At the negative electrode (anode)**: Hydrogen is oxidized to form water.
  - $2H_2(g) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$

The overall reaction is the conversion of hydrogen and oxygen to water.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

Advantages and Disadvantages of Using Hydrogen as a Fuel



Hydrogen is a possible alternative to fossil fuels in the future.

Electrolysis Review Questions

Gold-Plating and Electrolytes

When gold-plating a heat shield, the setup involves specific electrodes and an electrolyte:

- **Electrolyte**: Aqueous gold salt
- **Positive Electrode**: Gold
- **Negative Electrode**: Heat shield

Hydrogen-Oxygen Fuel Cell

In a hydrogen-oxygen fuel cell, electricity is generated by the reaction between **hydrogen** and **oxygen**.

Electrolysis of Binary Compounds

Identifying Products in Electrolysis

In the electrolysis of binary compounds in a molten state, specific products form at each electrode.

(a) Lead(II) Bromide

- Products formed: Lead and Bromine
- Symbol equation: $PbBr_2(l) \rightarrow Pb(l) + Br_2(g)$

(b) Sodium Chloride

- Products formed: Sodium and Chlorine
- Symbol equation: $2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)$

(c) Lithium Oxide

- Products formed: Lithium and Oxygen
- Symbol equation: $2Li_2O(l) \rightarrow 4Li(l) + O_2(g)$

(d) Iron(II) Chloride

- Products formed: Iron and Chlorine
- Symbol equation: $FeCl_2(l) \rightarrow Fe(l) + Cl_2(g)$

Experiment Analysis

Electrolyte

The electrolyte in the experiment shown in Figure 7.15 is **Lead(II) Chloride**.

Necessity of Molten Ionic Compound

The electrolyte must be a molten ionic compound because:

> Molten ionic compounds can conduct electricity due to the presence of mobile ions.

Products at Anode and Cathode

- **Anode**: Chlorine
- **Cathode**: Lead

Overall Equation

The overall equation for the electrolysis reaction is:

$PbCl_2(l) \rightarrow Pb(l) + Cl_2(g)$

Electrolytic Cell for Sodium Hydroxide Production

Identifying Gases

In the electrolytic cell used commercially to produce sodium hydroxide from concentrated aqueous sodium chloride:

- Gas X: Chlorine (Cl_2)
- Gas Y: Hydrogen (H_2)

Equations for Reactions at Electrodes

- Anode (producing Chlorine): $2Cl^- \rightarrow Cl_2 + 2e^-$
- Cathode (producing Hydrogen): $2H^+ + 2e^- \rightarrow H_2$

Porous Membrane

The porous membrane is present to:

> Separate the chlorine gas produced at the anode from the sodium hydroxide and hydrogen gas produced at the cathode, preventing them from mixing and reacting.

Property of Titanium

Since the anode used for this cell is titanium, a property of titanium is that it is **resistant to corrosion**.

Reduced X Collection with Iron Anode

If an iron anode is used, the amount of X (chlorine) collected is much reduced because:

> Iron is more reactive than titanium and will react with chlorine, reducing the amount of chlorine gas collected.

🔬 Copper Refining Apparatus

Ions in Aqueous Copper(II) Sulfate

The ions present in aqueous copper(II) sulfate are:

Cu^{2+} and SO_4^{2-}

Reaction at Cathode

The symbol equation for the reaction occurring at the cathode is:

$Cu^{2+} + 2e^- \rightarrow Cu$

Anode Mass Loss

(i) The anode decreased in mass because:

> Copper atoms from the impure copper anode are oxidized to copper ions (Cu^{2+}), which dissolve into the solution.

(ii) Calculating Percentage Purity:

- Loss of mass of the anode: 52g
- Gain in mass of the cathode: 48g
- Mass of copper dissolved: 52g
- Purity percentage: $\frac{48}{52} * 100$ = 92.3%

Observations Below the Anode

Just below the anode after electrolysis, you would observe:

> Accumulation of impurities, forming an "anode sludge" or "anode mud," which contains metals less reactive than copper.

Industrial Electrolytic Process

Identifying Anode and Cathode

- **Anode**: Iron
- **Cathode**: Chromium

Net Result of Electrolysis

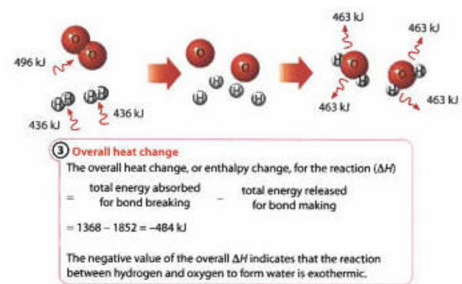
The net result of electrolysis is:

> Chromium from the aqueous chromium(II) sulfate solution is deposited onto the iron block, coating it with chromium.

Industrial Process Representation

This cell represents **electroplating**, specifically chromium plating of iron.

Industrial Process Setup



This figure illustrates the industrial process setup that we are analyzing. The process involves an external circuit connected to electrodes, with a pure copper cathode, an impure copper anode, and an aqueous copper(II) sulfate electrolyte.

Identifying the Anode

(a) In Figure 7.19, X is the anode. Explanation:

> The impure copper anode (X) dissolves into the electrolyte, providing copper ions for deposition at the cathode.

Industrial Application

(b) The industrial application of this process is **copper refining**.

Electron Flow Direction

(c) Electrons flow in rods X and Y towards end A (the cathode).

Laboratory Experiments

Ions in Aqueous Copper(II) Chloride

The ions present in aqueous copper(II) chloride are:

Cu^{2+} and Cl^-

Half-Equations for Reactions

(i) Set-up A (Carbon Electrodes):

- Anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
- Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

(ii) Set-up B (Copper Electrodes):

- Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
- Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Observations During Reaction

(i) Set-up A:

- Anode: Bubbles of chlorine gas are produced.
- Cathode: A deposit of copper metal forms on the electrode.
- Solution: The blue color of the copper(II) chloride solution fades.

(ii) Set-up B:

- Anode: The copper electrode dissolves, and the solution becomes more blue.
- Cathode: Copper is deposited on the electrode, increasing its mass.

Industrial Application

An industrial application that occurs in set-up B is **electroplating**, where a layer of copper is deposited onto another metal object.

Exothermic and Endothermic Reactions

Exothermic Reactions

An **exothermic reaction** transfers thermal energy to the surroundings, leading to an increase in the temperature of the surroundings. In an exothermic reaction, heat exits the reaction mixture, and the surroundings become warmer.

Examples of exothermic changes:

- Combustion of fuels
- Respiration
- Corrosion of metals (e.g., rusting of iron)
- Neutralization (reaction between acid and alkali)
- Dissolving of anhydrous salts (e.g., anhydrous copper(II) sulfate) in water

Endothermic Reactions

An **endothermic reaction** takes in thermal energy from the surroundings, leading to a decrease in the temperature of the surroundings. In an endothermic reaction, heat enters the reaction mixture, and the surroundings become cooler.

Examples of endothermic changes:

- Evaporation
- Thermal decomposition
- Photosynthesis
- Dissolving of ionic salts in water

Enthalpy Change (ΔH)

> The enthalpy change (ΔH) is the transfer of thermal energy during a reaction, measured in kilojoules (kJ).

- ΔH is negative for exothermic reactions (ΔH < 0).
- ΔH is positive for endothermic reactions (ΔH > 0).
- $\Delta H = \text{total energy of products} - \text{total energy of reactants}$

Worked Example 8A

When 1 mol of iron reacts with 1 mol of copper(II) sulfate in aqueous solution, 150 kJ of heat is given out. Write the equation for the reaction and state the enthalpy change.

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)\ \Delta H = -150\ \text{kJ}$

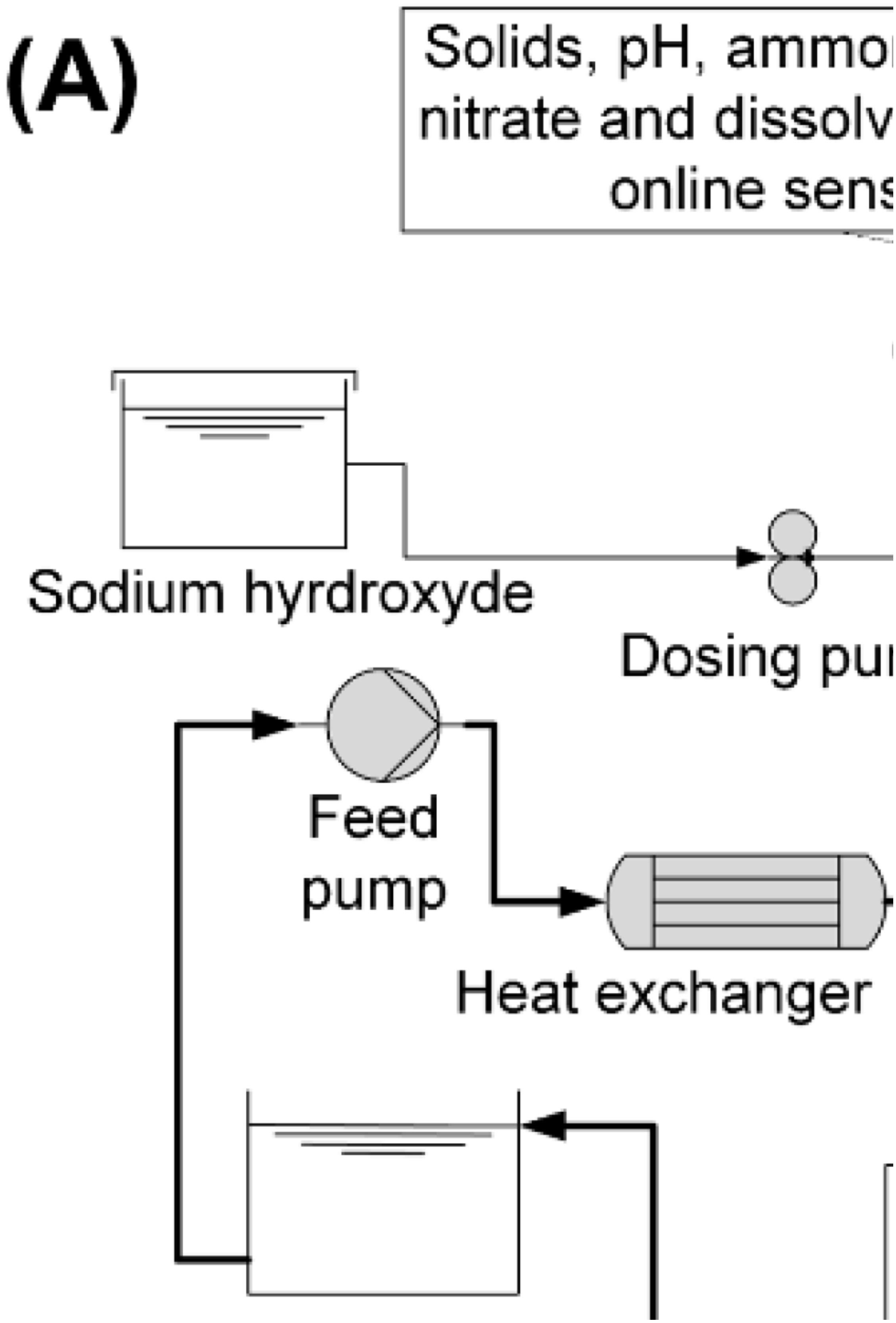
Worked Example 8B

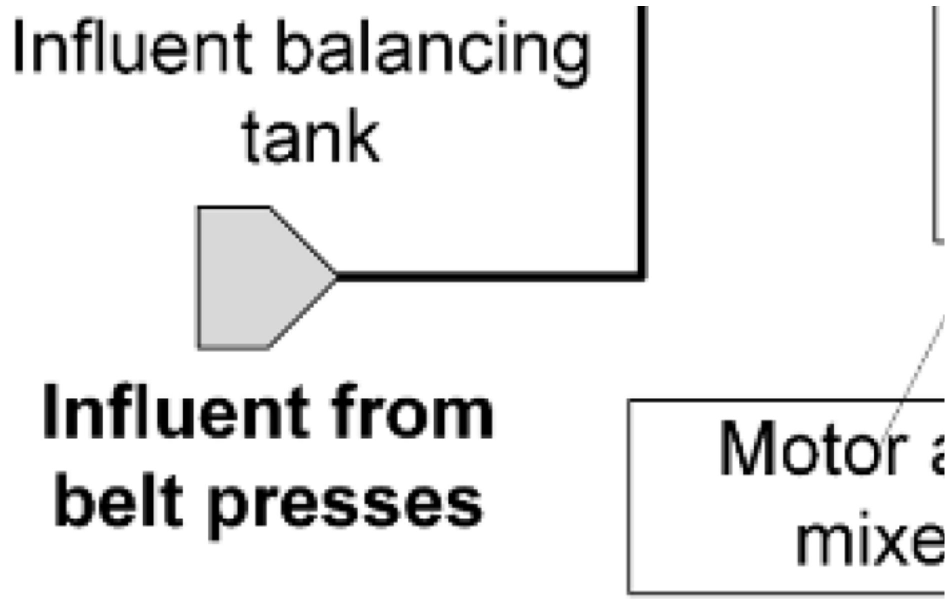
When 1 mol of solid ammonium chloride decomposes to form ammonia gas and hydrogen chloride gas, 136 kJ of heat energy is absorbed. Write the equation for the reaction and state the enthalpy change.

$NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)\ \Delta H = +136\ \text{kJ}$

Reaction Pathway Diagrams

Exothermic Reaction Pathway


(A)
Solids, pH, ammo
nitrate and dissolv
online sens
Sodium hyrdroxyde
Dosing pu
Feed
pump
Heat exchanger


Influent balancing
tank
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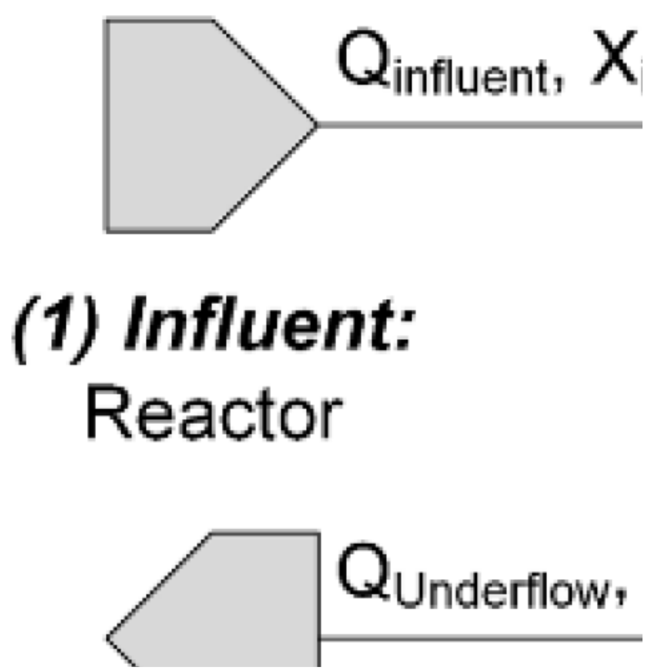
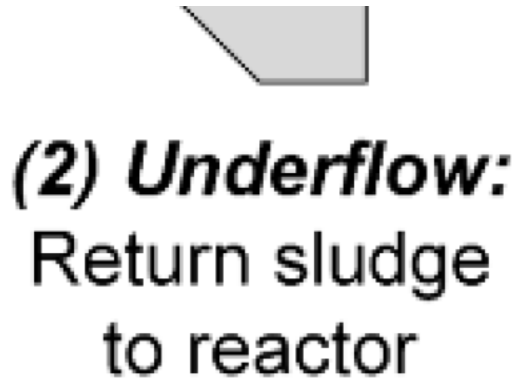
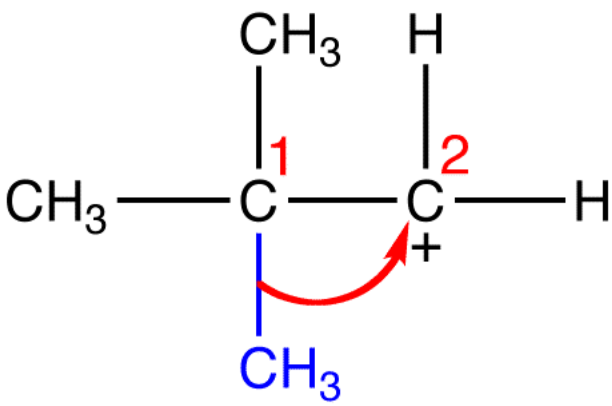
(B)


Q influent, X
(1) Influent:
Reactor
Q Underflow,



The reaction pathway diagram for an exothermic reaction shows that the total energy of the products is less than that of the reactants, as thermal energy is given out to the surroundings.

Endothermic Reaction Pathway



The reaction pathway diagram for an endothermic reaction shows that the total energy of the products is more than that of the reactants, as thermal energy is absorbed from the surroundings.

Activation Energy

> Activation energy (E_a) is the minimum energy that colliding particles must have to react.

A reaction will occur only when reactant particles possess energy equal to or more than the activation energy.

Exothermic Reaction Pathway with Activation Energy

Exothermic Activation Energy

Endothermic Reaction Pathway with Activation Energy

Endothermic Activation Energy

The activation energy acts as an energy barrier that must be overcome by the reactants before they can react to form products.

Bond Breaking and Bond Making

Energy Changes in Reactions

All reactions involve the breaking of bonds or the forming of new bonds. Energy changes in reactions are caused by the making and breaking of chemical bonds.

- Bond breaking is an **endothermic** process (energy is absorbed).
- Bond making is an **exothermic** process (energy is given out).

Reaction Exothermic or Endothermic

The overall enthalpy change determines whether a reaction is exothermic or endothermic:

$\Delta H = \text{total energy absorbed for bond breaking} - \text{total energy released for bond making}$

- Exothermic reaction: total energy absorbed for bond breaking < total energy released for bond making
- Endothermic reaction: total energy absorbed for bond breaking > total energy released for bond making

Calculating Enthalpy Change Using Bond Energies

> Bond energy is the energy required to break a covalent bond, and the same amount of thermal energy is released whenever the same covalent bond is formed.

Thermodynamics

Instant Coolers and Phase Transitions

Instant car coolers utilize the principle of **latent heat of vaporization**. When the cooler is sprayed, a liquid with a low boiling point vaporizes inside the overheated car. This vaporization process requires energy, which is absorbed from the surroundings, causing the temperature of the car to drop rapidly.

Combustion vs. Rusting

Similarities:

- Both are **exothermic** reactions.
- Both involve **oxidation**.

Differences:

- Combustion is **fast**, rusting is **slow**.
- Combustion involves a **flame**, rusting does not.
- Combustion involves **hydrocarbons**, rusting involves **iron**.

Reaction Pathway Diagrams

Exothermic or Endothermic?

Based on Figure 8.14 (not provided), determine whether the reaction between compound Q and dilute hydrochloric acid is exothermic or endothermic. An **exothermic reaction** releases heat, while an **endothermic reaction** absorbs heat. The diagram should show whether the energy of the products is lower (exothermic) or higher (endothermic) than the energy of the reactants.

(a) The reaction is **exothermic** because the energy of the products is lower than that of the reactants, thus energy is released in the reaction.

Temperature Change

(b) The temperature of the mixture will **rise** when solid Q is added to dilute hydrochloric acid.

(c) Because the reaction is exothermic, it releases heat into the mixture, causing the temperature to rise.

(d) Compound Q is likely a carbonate, such as $CaCO_3$, which reacts with hydrochloric acid to produce carbon dioxide gas.

Hydrogen and Oxygen Reaction

Energy Comparison

(a) The mixture of hydrogen and oxygen has a greater amount of energy than water. This is because the reaction between hydrogen and oxygen to form water is exothermic, meaning energy is released during the formation of water.

Reaction Pathway Diagram

(b) Draw a reaction pathway diagram for the reaction between hydrogen and oxygen to form water, labeling the activation energy (E_a) and the enthalpy change (ΔH).

Bond Energies

(c) If the energy released when forming the bonds of 2 moles of water is 1852 kJ, then the energy needed for breaking the bonds of 2 moles of hydrogen and 1 mole of oxygen can be calculated using the principle of conservation of energy.

$$Energy_{breaking} = Energy_{forming} = 1852, kJ$$

Titration Experiments

Temperature Rise Comparison

Experiments 1 and 4 will produce the same temperature rise.

- In Experiment 1: 50 cm³ of 1.0 mol/dm³ HNO_3 and 50 cm³ of 1.0 mol/dm³ $NaOH$.
- In Experiment 4: 200 cm³ of 2.0 mol/dm³ HCl and 200 cm³ of 2.0 mol/dm³ KOH.

Both experiments have the same number of moles of acid and base reacting (0.05 moles).

Ammonia and Chlorine Reaction

Given the reaction:

$$2NH_3(g) + 3Cl_2(g) \rightarrow N_2(g) + 6HCl(g)$$

Bond Analysis

(a) (i) Bonds broken:

- 6 N-H bonds
- 3 Cl-Cl bonds

(a) (ii) Energy needed to break bonds:

$$E_{broken} = (6 \times 390) + (3 \times 242) = 2340 + 726 = 3066, kJ$$

(b) (i) Bonds formed:

- 1 N≡N bond
- 6 H-Cl bonds

(b) (ii) Energy released when bonds are formed:

$E_{formed} = (1 \times 945) + (6 \times 431) = 945 + 2586 = 3531, kJ$

(c) The reaction is **exothermic** because the energy released when the bonds are formed (3531 kJ) is greater than the energy needed to break the bonds (3066 kJ).

(d) Energy change when 1.0 mole of ammonia reacts:

First, we have to half the energy needed to break the bonds and the energy released when bonds are formed as the equation gives the energy needed for *2* moles of ammonia and we only want it for *1* mole.

$E_{broken} = 3066/2 = 1533, kJ$

$E_{formed} = 3531/2 = 1765.5, kJ$

$\Delta H = E_{formed} - E_{broken} = 1765.5 - 1533 = -232.5, kJ$

Rate of Reaction

Slowing Down Milk Souring

Two methods to slow down milk souring:

- **Refrigeration:** Lowers the temperature, slowing down the bacterial activity that causes souring.
- **Pasteurization:** Heats the milk to kill bacteria, extending its shelf life.

Chemists need to know how fast a reaction proceeds to optimize industrial processes, ensure safety, and understand chemical phenomena.

Reaction Rate Comparison

Arranging reactions from slowest to fastest:

1. Weathering of stone sculptures
2. Rusting of iron
3. Ripening of bananas
4. Fermentation of grapes
5. Explosions

Physical and Chemical Changes

Physical Change

> A physical change alters the form of a substance but does not result in the formation of new substances.

Examples: mixing iron with sulfur, dissolving salt in water, melting ice.

Chemical Change

> A chemical change forms one or more new substances.

Examples: rusting of iron nails, decomposition of substances, explosion of fireworks.

Factors Affecting the Rate of Reaction

Factors Influencing Reaction Rate

- Concentration of solutions
- Pressure of gases
- Surface area of solids
- Temperature
- Catalysts

Concentration

Increasing the concentration of a reactant increases the rate of reaction.

Pressure

Increasing the pressure of a gaseous reactant increases the rate of reaction.

Surface Area

Increasing the surface area of a solid reactant increases the rate of reaction.

Temperature

Increasing the temperature of a reactant increases the rate of reaction.

Catalysts

A catalyst increases the rate of a chemical reaction and remains chemically unchanged at the end of the reaction.

Investigating the Rate of Reaction in the Laboratory

Measuring Gas Volume

The rate of reaction can be determined by measuring the volume of gas produced in a certain time.

$$Rate, of, reaction = \frac{volume, of, gas, produced, (cm^3)}{time, taken, (s)}$$

Measuring Mass Change

The rate of reaction can be determined by measuring the change in mass of a reaction mixture in a certain time.

Collision Theory

> The collision theory states that for a reaction to occur, reactant particles must collide with each other with sufficient energy (activation energy).

Collision Theory Factors

- Number of particles per unit volume
- Frequency of collisions between particles
- Kinetic energy of particles
- Activation energy

Catalyst Effect

A catalyst decreases the activation energy of a reaction, thereby increasing the rate of reaction.

Collision Theory and Effective Collisions

For a reaction to occur, reactant particles must collide with:

1. Sufficient **kinetic energy**.
2. The correct **orientation**.

The kinetic energy must be equal to or greater than the **activation energy**. When these conditions are met, **effective collisions** occur, leading to the formation of product particles.

> **Effective Collisions:** Collisions between reactant particles that result in the formation of product particles.

Factors Affecting Reaction Rate Explained by Collision Theory

The **frequency of effective collisions** directly impacts the reaction rate. An increased frequency leads to a higher reaction rate.

> **Frequency:** The number of times something occurs within a given period.

Concentration of Reactants

Increasing the concentration of a reactant increases the reaction rate.

- **Lower concentration:** Reactant particles are spread far apart.
- **Higher concentration:** More reactant particles occupy the same volume.

With a higher concentration, collisions between reactant particles become more frequent, increasing the frequency of effective collisions and, consequently, the reaction rate.

Imagine a theme park with bumper cars. The greater the number of bumper cars, the more frequently they will collide with each other.

Pressure of Gaseous Reactants

Increasing the pressure of a gaseous reactant also increases the reaction rate.

- **Lower pressure:** Gaseous reactant particles are spread far apart.
- **Higher pressure:** Gaseous reactant particles are closer together, compressed into a smaller volume.

Increasing pressure is similar to increasing concentration. Higher pressure leads to more frequent collisions, resulting in a higher frequency of effective collisions and an increased reaction rate.

Surface Area of Solid Reactants

If a solid reactant is broken into smaller pieces, its total surface area increases.

- More surfaces are exposed, allowing for more collisions to occur.

Consider a solid reactant with dimensions:

- (a) 2cm x 1cm x 0.5cm, with a surface area = $2(2x0.5) + 2(2x1) + 2(1x0.5) = 7cm^2$
- (b) Broken into 8 pieces of 1cm x 0.5cm x 0.5cm, with a surface area = $8 * 2(0.5x0.5) + 8 * 4(1x0.5) = 21cm^2$

The smaller the size of the particles, the larger the total surface area.