



Chapter 1

Background Knowledge

Chapter Objectives

1. To understand standard atomic theory
 2. To define elements, compounds and mixtures
 3. To recall the properties of states
 4. To distinguish between physical and chemical properties and between physical and chemical changes
 5. To perform calculations using metric units, scientific notation and unit conversions
 6. To recall the measurement of mass, volume and density
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Matter

Matter is made up of atoms, which are the smallest objects which cannot be further decomposed chemically. The presence of atoms was originally hypothesized due to the discovery of several fundamental chemical laws and three are important to the SAT test, as discussed below:

(i) The law of conservation of mass

The law of conservation of mass states that mass can neither be created nor destroyed. In other words, the mass of a system is conserved in a chemical reaction.

(ii) The law of definite proportion

Another one is known as *the law of definite proportion*. It states that a compound always contains the same proportion of elements by mass. For example:

18 g of water contains 2 g of hydrogen and 16 g of oxygen. The mass ratio of hydrogen to oxygen is therefore 1 : 8. In 36 g of water, there are 4 g of hydrogen and 32 g of oxygen. The mass ratio of hydrogen to oxygen is still the same as 1 : 8. Thus, the mass ratio is maintained.

(iii) The law of multiple proportions

The third one is *the law of multiple proportions* which states that when a series of compounds is formed from two elements, the ratios of the masses of one element combining with one gram of another element in each compound can always be simplified to small whole numbers. For example:

Carbon and oxygen can form two different compounds, compound X and compound Y. Compound X is produced by combining one gram of carbon and 1.33 grams of oxygen, while compound Y is produced by combining one gram of carbon and 2.66 grams of oxygen. The ratio of oxygen required to react with one gram of carbon in compound X to that in compound Y is thus 1.33 : 2.66, which can be reduced to 1 : 2, a whole number ratio.

Dalton's Atomic Theory

The properties of atoms are summarized by *Dalton's atomic theory*, which proposed the followings:

1. Atoms of the same kind comprise an *element*.
2. Atoms of an element possess the same chemical properties.
3. When atoms of different elements combine together, a *compound* is formed. A given compound always has a fixed ratio of different elements.
4. A chemical reaction involves rearrangement of atoms only. Atoms can never be destroyed, created or changed in a chemical reaction.

Elements, Compounds and Mixtures

Matter can be classified according to its composition. *Pure substances* contain only one kind of element or compound. For example, graphite only contains carbon atoms; water is a compound of hydrogen and oxygen.

On the other hand, *impure substances* contain two or more kinds of pure substances physically mixed together and they are called *mixtures*. Sea water consists of salts and water; while air contains nitrogen gas, oxygen gas and other gases; so both of them are mixtures.

Mixtures can be either *homogeneous* or *heterogeneous*. Homogenous mixtures contain substances which are distributed evenly throughout the mixture. For example, the composition of salt and water in salt solution is the same everywhere in the solution. Heterogeneous mixtures contain substances that cannot be mixed thoroughly, like sand and water.

The characteristics of elements, compounds and mixtures are summarized in Table 1.1.

Table 1.1 Characteristics of Elements, Compounds and Mixtures

Matter	Pure substances		Impure substances
	Elements	Compounds	Mixtures
Composition	Contain only one kind of atom.	Contain two or more kinds of atoms chemically bonded together; have a definite composition of elements.	Contain two or more kind of pure substances, physically mixed together; can have various ratios of constituent substances.
Properties	Possess distinct properties for different elements.	Possess properties that are entirely different from the constituent elements.	Retain the properties of the constituent substances.
Separation	Cannot be further decomposed into anything simpler by chemical means	Can be decomposed back to elements by chemical means like electrolysis	Can be separated into constituent substances by physical means like distillation

Solids, Liquids and Gases

Matter can be categorized into three states. **Solids** have a fixed shape and volume; **liquids** have the shape of a container and a fixed volume; **gases** have both the shape and volume of a container. At a microscopic level, all matter can be regarded as an aggregation of tiny units called particles. The particles in solids are closely packed together and they can only vibrate about a fixed position due to strong attractive forces. In liquids, particles are still closely packed but they can move around each other randomly to some extent since the attractive force is weaker. In gases, particles are separated far away from each other and they are essentially free to move fast and randomly within a container. There is almost no attractive force between the particles in gases.

The properties of gases, liquids and solids are summarized in Table 1.2 on the next page.

Table 1.2 Properties of Gases, Liquids and Solids

	Gases	Liquids	Solids
Shape	Of the container	Of the container	Fixed
Volume	Of the container	Fixed	Fixed
Compressibility	High	Negligibly low	Negligibly low
Viscosity	Low	Medium to High	Infinitely high
Arrangement of particles	Random	Random	Packed in a lattice
Separation of particles	Separated far apart	Closely packed	Closely packed
Attractive forces between particles	Negligible at low pressure	Strong enough to hold particles close together but not fixed in space	Strong to hold particles in fixed positions
Motion of particles	Fast, random motion	Slow, random motion around other particles	Vibration about fixed positions

Physical Properties and Chemical Properties

Different substances can be distinguished by their *physical properties* and *chemical properties*.

Physical properties are the characteristics of a substance which do not involve a change in composition. It includes physical state, appearance, density, hardness, melting point and boiling point.

Chemical properties are the ability of a substance to react with other substances or decompose on its own to form new substances. For example, carbon burns with oxygen gas to form carbon dioxide. Water can be decomposed by electricity to form hydrogen gas and oxygen gas.



Physical Changes and Chemical Changes

The changes of matter can either be *physical changes* or *chemical changes*.

Physical changes only alter the physical properties of a matter but not its composition, such as the breaking of glass, magnetization of iron and melting of ice. Physical changes are usually easy to reverse. For example, water can be changed back to ice if it is put into a freezer.

Chemical changes involve the formation of new substances. They are usually difficult to reverse. Examples are iron rusting and paper burning.

Measurements and Unit Conversion

SI system

Measurements must be recorded in a systematic way such that results from different experiments can be compared. Number itself is not sufficient to specify an absolute measurement because instruments based on different standards may be used. Scientists thus define certain standards as units. In the SAT exam, the *metric system* is used. Particularly, units from the *SI system*, an internationally well-recognized metric system, are often used. The SI system defines seven fundamental SI units, of which six are commonly applied in chemistry. They are shown in Table 1.3 below.

Table 1.3 The Fundamental SI Units

Physical quantity	Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s
Temperature	Kelvin	K
Electric current	Ampere	A
Amount of substance	Mole	mol

Sometimes, the fundamental SI units are not convenient for describing very large or very small quantities. **Prefixes** are added to change the order of magnitude of a unit. The prefixes most encountered in SAT chemistry are listed in Table 1.4. Examples applying the prefixes are given in Table 1.5 on the next page.

Table 1.4 Prefixes Used in the SI System

Prefix	Abbreviation	Multiple	Exponential Notation
mega-	M	1 000 000	10^6
kilo-	k	1 000	10^3
deci-	d	0.1	10^{-1}
centi-	c	0.01	10^{-2}
milli-	m	0.001	10^{-3}

Table 1.5 Examples of Common Prefixed Units

Physical quantities	Relationship between quantities	Abbreviation
kilograms and grams	1 kilogram = 1000 grams	1 kg = 1000 g
liters and milliliters	1 liter = 1000 milliliters 1 milliliter = 0.001 liters	1 L = 1000 mL 1 mL = 0.001 L
meters and nanometers	1 meter = 1 000 000 000 nanometers 1 nanometer = 0.000 000 001 meters	1 m = 1 000 000 000 nm 1 nm = 0.000 000 001 m

Scientific Notation

The other way to amend the problem of large or small numbers is to use **scientific notation**. A very large number or a very small number is written as a product of a number, which has a value between 1 inclusive and 10 exclusive, and powers of ten.



Examples: $987000000 \text{ m} = 9.87 \times 10^8 \text{ m}$
 $0.00000234 \text{ kg} = 2.34 \times 10^{-6} \text{ kg}$

The steps of writing scientific notation are as follows.

1. Put a new decimal point after the first non-zero digit of the number.
2. Count the number of places that the decimal point needs to move from the original place to the new one as the value of the power of ten.
3. If the decimal point moves to the left, the power of ten is positive. If it moves to the right, the power of ten is negative.
4. Write the non-zero digits with the new decimal point in the front, followed by the power of ten.

Examples are given in Table 1.6.

Table 1.6 Examples of writing scientific notation

	Examples	Workings	Scientific notation
Number much greater than 1	987000000 m	$9.87000000.$ 	$9.87 \times 10^8 \text{ m}$
Number much smaller than 1	0.00000234 kg	$0.000002.34$ 	$2.34 \times 10^{-6} \text{ kg}$

Unit conversion

Sometimes, units of a particular physical quantity are not related to each other by a power of ten and thus are more difficult to convert between each other. For example, 1 hour equals 60 minutes. We can apply a technique called **dimensional analysis** to convert between the units. Say, we would like to find out how many minutes are the same as 2.5 hours. As we know, 1 hour equals 60 minutes, so we multiply the hours by a conversion factor. The conversion factor is a ratio of the two sides of the equivalence statement, which is arranged such that the unwanted unit can be cancelled out while the desired unit remains. The working is shown below.

$$1 \text{ h} = 60 \text{ min}$$

$$2.5 \text{ h} = 2.5 \text{ h} \times \frac{60 \text{ min}}{1 \text{ h}} = 150 \text{ min}$$

This technique is particularly useful in stoichiometric calculations, in which enormous unit conversions are required.

Mass

Mass is the amount of substance. 36 g of hydrogen contains twice as many atoms as 18 g of hydrogen. In contrast to physics, mass and weight are used interchangeably in chemistry by convention. Mass of a substance is usually measured by a weighing balance. The kilogram is the fundamental SI unit of mass. However, on a laboratory scale, the gram is used more often. Milligrams are sometimes used to give the amount of a substance of tiny mass. The conversion between these units is shown below:

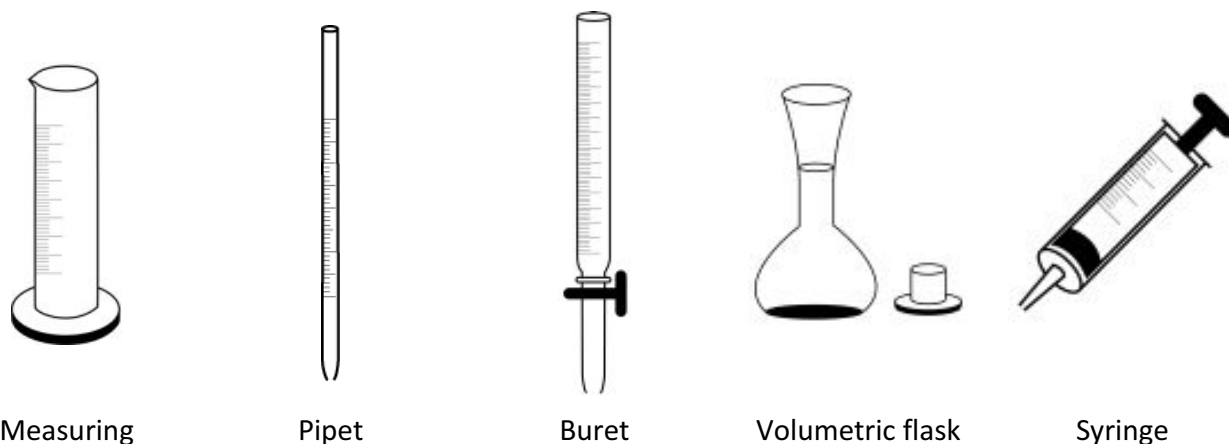
$$1 \text{ kg} = 1000 \text{ g}$$
$$1 \text{ g} = 1000 \text{ mg}$$

Scientific notation is often applied when the mass of an atom is expressed. For example, one hydrogen atom weighs approximately 1.67×10^{-19} kg.

Volume

Volume shows the amount of space a substance occupies. The volume of an insoluble solid is determined by measuring the volume of a liquid displaced when the solid is immersed in a container filled fully with the liquid. The volume of liquids is usually measured with glassware. A **measuring cylinder** is used to estimate the volume of a liquid roughly. A **pipet** can deliver a fixed and accurate volume of liquid. A **buret** can measure the amount of liquid added to glassware accurately. A **volumetric flask** allows a solution to be prepared with a fixed and accurate volume. For a gas, its volume is usually measured with a **graduated syringe**. The apparatus are shown in Figure 1.1.

Figure 1.1 Apparatus Used for Determination of Volume



cylinder

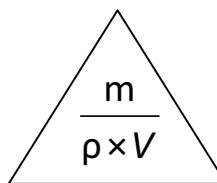
Although the fundamental SI unit for volume is m^3 , chemists usually use cm^3 , dm^3 , mL or L to describe volumes of liquid and gas in a laboratory scale. The relationships among these units are as below:

$1 \text{ m}^3 = 1000 \text{ dm}^3$	$1 \text{ dm}^3 = 1 \text{ L}$
$1 \text{ dm}^3 = 1000 \text{ cm}^3$	$1 \text{ cm}^3 = 1 \text{ mL}$
$1 \text{ L} = 1000 \text{ mL}$	

Density

Density is a measure of the mass of a substance per unit volume. It can be represented by the following formula, where ρ denotes density; m denotes mass and V denotes volume.

$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$
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Once a substance has its mass and volume known, its density can be determined with the formula. The common units for density are g/cm^3 , g/mL and g/L . The relationship between the units is:

$1 \text{ g/cm}^3 = 1 \text{ g/mL}$
$1 \text{ g/mL} = 1000 \text{ g/L}$

The triangle on the right of the formula provides a convenient way to determine the unknown once the others are known. The unknown quantity is equal to the expression in the triangle with the unknown covered by a finger. Suppose we want to find the density of copper and we know that 178 g of copper has a volume of 20 cm^3 . By covering ρ with a finger, the expression for ρ is equal to m divided by V which gives:

$$\rho = \frac{m}{V} = \frac{178}{20} = 8.9 \text{ g/cm}^3$$

By the same technique, the mass of 100-mL water, with a density of 1 g/mL at $4 \text{ }^\circ\text{C}$, can be determined. The triangle notation is used with various formulas throughout the text.

$$m = \rho \times V = 1 \times 100 = 100 \text{ g}$$

If two substances have the same volume but different densities, then the one with a greater density is heavier. Therefore, 1 cm^3 of lead is heavier than 1 cm^3 of water.

If two substances have the same mass but different densities, then the one with a greater density is smaller in size. Therefore, 1 g of lead occupies a smaller volume than 1 g of water.

Density can also tell whether a substance floats or sinks in a fluid (i.e. a liquid or a gas). Ice has a density smaller than water and it floats on water. On the other hand, metal has a greater density than water and it sinks.

The densities of solid and liquid are often approximated as constants because the volumes of solid and liquid do not change significantly with respect to temperature (and the mass of a substance is always conserved). However, the volume of gas varies greatly with temperature. The density of gas, therefore, changes with temperature. The relationship between temperature and density of gas will be discussed in Chapter 5.

Summary

1. Atoms cannot be destroyed or created in chemical reactions.
 2. An element contains one kind of atom; a compound contains a fixed ratio of elements; a mixture contains pure substances mixed together.
 3. Solid has a fixed shape and volume; liquid has the shape of a container and fixed volume; gas has both the shape and volume of a container.
 4. Physical properties refer to the characteristics of a substance related to physical changes, which do not alter the composition of the substance; chemical properties refer to the ability of a substance to undergo chemical changes, which involve formation of new substances.
 5. Units are required to make measurements meaningful. Both SI units and scientific notation are tools to describe physical quantities in a systematic and convenient way. Dimensional analysis is a useful method to convert quantities between different units.
 6. Mass, volume and density are often measured in experiments. Different apparatus is used to measure these quantities of a substance in different states.
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Key Words

The law of conservation of mass
The law of definite proportion
The law of multiple proportions
Dalton's atomic theory
Element
Compound
Pure substance
Impure substance
Mixture
Homogeneous
Heterogeneous
Solid
Liquid
Gas
Physical properties

Chemical properties
Physical changes
Chemical changes
Metric system
SI system
Scientific notation
Dimensional analysis
Mass
Volume
Measuring cylinder
Pipet
Buret
Volumetric flask
Graduated syringe
Density

Review Questions

Part A

Questions 1-4 refer to the following.

- (A) Mass
- (B) Density
- (C) Volume
- (D) Boiling point
- (E) Flammability

1. Is a chemical property
2. Has as its fundamental unit the kelvin
3. Is a measure of mass per unit volume
4. Is independent of the external conditions for all substances

Part B

- | | I | II |
|------|---|--|
| 101. | The total mass of products obtained is the same as the total initial mass of the reactants in a complete reaction | BECAUSE the law of definite proportion states that the ratio of elements in a compound is fixed. |
| 102. | Air is a compound | BECAUSE air contains almost a fixed composition of nitrogen gas and oxygen gas. |
| 103. | 100 grams of lead is heavier than 100 grams of graphite | BECAUSE the density of lead is greater than that of graphite. |

Part C

5. Which of the following statements CANNOT be deduced from Dalton's atomic theory?
- (A) An element contains only one kind of atoms.
 - (B) A compound always has a fixed ratio of different kinds of atoms.
 - (C) There is no loss in the total mass in a chemical reaction.
 - (D) Gas contains atoms which are in constant and random motion.
 - (E) All atoms of hydrogen can react with oxygen to form water.
6. Which of the following properties can be attributed to carbon dioxide?
- I. It exhibits the chemical properties of both carbon and oxygen.
 - II. It can be produced through the chemical reaction between carbon and oxygen.
 - III. It cannot be decomposed by physical means.
- (A) I only
 - (B) II only
 - (C) I and II only
 - (D) II and III only
 - (E) I, II and III
7. Which of the following is equivalent to one cubic decimeter?
- I. $1 \times 10^{-3} \text{ m}^3$
 - II. $1 \times 10^3 \text{ mL}$
 - III. $1 \times 10^6 \text{ cm}^3$
- (A) I only
 - (B) II only
 - (C) I and II only
 - (D) I and III only
 - (E) I, II and III
8. What is the density of metal X if 5.0 cm^3 of the metal weighs 9.0 g?
- (A) $1.8 \times 10^{-3} \text{ g/cm}^3$
 - (B) $5.6 \times 10^{-1} \text{ g/cm}^3$
 - (C) $1.8 \times 10^3 \text{ g/cm}^3$
 - (D) $5.6 \times 10^{-4} \text{ g/L}$
 - (E) $1.8 \times 10^3 \text{ g/L}$
9. If solid A and B are of equal volume and solid A has twice the density of solid B, which of the following is true?
- (A) Solid A is two times heavier than half of solid B.
 - (B) Solids A and B are of equal mass
 - (C) The mass of solid B is the same as half of solid A.
 - (D) Solid B is two times heavier than half of solid A.
 - (E) Solid B is two times heavier than solid A.

Answers and Explanation

Question number	Answer	Explanation
1	E	Flammability of a substance indicates whether a substance reacts with oxygen in air. Therefore, flammability involves formation of new substances.
2	D	Boiling point is a measure of temperature at which a solid turns into a liquid. The fundamental SI unit of temperature is kelvin.
3	B	By definition, density of a substance is the mass of a substance per unit volume.
4	A	Mass is a measure of the amount of substance. It does not change by external factors, like varying temperature, pressure, etc.
101	T, T	Statement I is correct because mass is conserved in a chemical reaction. Statement II is correct because it is the correct definition of the law of definite proportion. However, statement II cannot explain statement I as they are not related to each other.
102	F, T	Air contains nitrogen gas, oxygen gas and other gases, which are physically mixed together. Therefore, statement I is incorrect. However, air is a homogeneous mixture which means its composition is the same from region to region. Therefore, statement II is correct.
103	F, T	Statement I is incorrect because substance with the same mass always weigh the same. Statement II is correct because lead is a dense metal whose density is greater than most non-metal.
5	D	Dalton's atomic theory does not mention the motion of atoms.
6	D	Carbon dioxide is a compound of carbon and oxygen. It does not have the properties of both constituent elements and can only be decomposed by chemical means, but not physical means.

Question number	Answer	Explanation
7	C	$1 \text{ dm}^3 = 1 \text{ dm}^3 \times \frac{1 \text{ m}^3}{1000 \text{ dm}^3} = 1 \times 10^{-3} \text{ m}^3$ $1 \text{ dm}^3 = 1 \text{ dm}^3 \times \frac{1000 \text{ cm}^3}{1 \text{ dm}^3} = 1 \times 10^3 \text{ cm}^3$ $1 \text{ dm}^3 = 1 \times 10^3 \text{ cm}^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} = 1 \times 10^3 \text{ mL}$
8	E	$\rho = \frac{m}{V} = \frac{9.0 \text{ g}}{5.0 \text{ cm}^3} = 1.8 \text{ g/cm}^3$ $1.8 \text{ g/cm}^3 = \frac{1.8 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \times 10^3 \text{ cm}^3}{1 \text{ L}} = 1.8 \times 10^3 \text{ g/L}$
9	C	<p>If volume is constant, mass is directly proportional to density. Solid A has twice the density of solid B, so solid A is two times heavier than solid B. In other words, half of solid A is of the same mass as solid B.</p>