MINISTRY OF EDUCATION AND TECHNICAL EDUCATION

وزارة التربية والتعليم والتعليم الفني

Chemistry

Egyptian Baccalaureate

2^{nd} Year

Part One

11

SPRIX

Chemistry

Egyptian Baccalaureate

2nd Year

Advised by

INTERNATIONAL BACCALAUREATE®

The International Baccalaureate™

has reviewed the pedagogical and assessment approaches underpinning this textbook.

Ministry of Education and Technical Education

New Administrative Capital

Cairo, Egypt

Introduction

Chemistry helps us understand the substances that make up our world, how they behave, and how they change. From the air we breathe and the materials we use to the medicines we take, the energy we produce, and the technologies that shape modern life, chemistry is part of almost everything around us.

This Grade 11 Chemistry textbook has been developed in alignment with the vision of the Egyptian Baccalaureate System to provide students with a rigorous, engaging, and meaningful learning experience. It is designed not only to build scientific knowledge, but also to develop the ability to question, investigate, analyze evidence, solve problems, and apply learning in new and unfamiliar situations.

Throughout this course, you will explore two major areas of chemistry: States of Matter and Change and Equilibrium of Matter. You will investigate how matter changes from one state to another, how gases behave, how solutions and crystals are structured, and how physical properties can be explained using scientific principles and models. You will also examine chemical reactions, energy changes, electrochemical cells, electrolysis, reaction rates, chemical equilibrium, ionization, aqueous solutions, and solubility.

These topics are closely connected. Understanding the behavior of gases and solutions, for example, helps explain many natural and industrial processes. Studying enthalpy allows us to understand the energy involved in chemical reactions. Electrochemistry connects chemical change with the production and use of electrical energy, while reaction rates and equilibrium help explain why reactions occur as they do and how their conditions can be controlled.
Each lesson is designed to take you through a clear learning journey. You will begin by connecting new ideas to prior knowledge, familiar observations, or real-world phenomena. You will then explore scientific concepts through questions, examples, investigations, models, data, and problem-solving activities. As your understanding develops, you will have opportunities to practise what you have learned and apply it independently.

Chemistry is not learned through memorization alone. Throughout this textbook, you will be asked to observe carefully, identify patterns, interpret evidence, make predictions, construct explanations, perform calculations, and justify your conclusions. Practical investigations and scientific reasoning will help you understand not only what happens during a chemical or physical process, but also why it happens.

The textbook also develops essential scientific skills, including measurement, classification, prediction, inference, experimentation, data interpretation, modelling, problem-solving, and scientific communication. These skills will help you approach questions in the way scientists do: by examining evidence, testing ideas, evaluating explanations, and refining conclusions. Assessment is integrated throughout the learning process. Questions and activities will allow you to check your understanding, practise key concepts, and prepare for more demanding applications of knowledge. You will encounter tasks that require recall and understanding, but also analysis, reasoning, calculation, interpretation, and the application of chemistry to unfamiliar situations. Where appropriate, the course connects chemistry to real-world contexts such as energy, industry, health, technology, materials, and environmental sustainability. These applications show how chemical knowledge contributes to solving practical problems and addressing challenges faced by communities in Egypt and around the world.

You will also be encouraged to think about the Nature of Science: how scientific knowledge is developed, tested, communicated, and refined over time. Scientific ideas do not emerge in isolation. They are built through observation, experimentation, evidence, debate, and the contributions of scientists across different cultures and generations.

Most importantly, this textbook is designed to help you become an active participant in your own learning. You will be encouraged to ask questions, reflect on your understanding, learn from mistakes, and make connections between different ideas. The aim is not simply to prepare you for an examination, but to develop the scientific knowledge, habits of mind, and confidence needed for future study and for life in an increasingly scientific and technological world.

As you begin this journey, approach chemistry with curiosity. Look beyond equations and definitions to the ideas they represent. Ask why substances behave as they do, examine the evidence carefully, and use what you learn to make sense of the world around you.
Chemistry is the study of matter and change, but learning chemistry can also change the way you see the world.

How to use this book

Every lesson follows a scientific learning journey:
Observe ▪ Question ▪ Investigate ▪ Explain ▪ Apply ▪ Evaluate

1 **Guiding Question**
Each lesson begins with a guiding question that focuses your thinking and helps you connect chemistry to real-world phenomena.

2 **Engaging Introduction**
A short scenario, phenomenon, or authentic context introduces the lesson and demonstrates why the topic matters in everyday life, technology, and society.

3 **Point!**
Key concepts, definitions, laws, principles, and scientific models are presented clearly and systematically.

4 **Exploration**
Investigate patterns, analyze evidence, identify relationships, and make predictions before formal explanations are introduced.

5 **Warm Up**
Questions and Examples that help you in applying knowledge and prepare for deeper learning.

6 **Explanation**
Step-by-step explanations demonstrate scientific reasoning, calculations, and problem-solving strategies.

7 **Think Like a Chemist**
Apply scientific thinking to justify conclusions, evaluate claims, and interpret experimental results.

8 **Think as an Engineer**
Use chemistry concepts to solve practical problems and make evidence-based decisions in technological and engineering contexts.

9 **Life Application**
Discover how chemistry influences sustainability, health, energy production, industry, transportation, and everyday life.

10 **Structure of Science**
Explore the scientific concepts, models, and principles that connect the lesson to broader scientific understanding.

11 **Nature of Science**
Examine how scientists collect evidence, design investigations, evaluate data, and build reliable explanations.

12 **Misconceptions Alert**
Identify frequent mistakes and misconceptions so that your understanding becomes more accurate and scientifically sound.

13 **Key equations at a glance**
Important formulas, symbols, and relationships are summarized for quick reference.

14 **Try**
Practice your learning
Apply new concepts through guided questions and structured exercises.

15 **Exercise**
Strengthen your understanding through increasingly challenging questions that require reasoning, calculations, and analysis.

16 **Challenge Your Thinking**
Tackle higher-order questions that require evaluation, synthesis, problem-solving, and scientific argumentation.

17 **Reflect and Grow**
Revisit key concepts from current and previous lessons to strengthen long-term understanding and exam readiness.

Your lessons and the EB syllabus

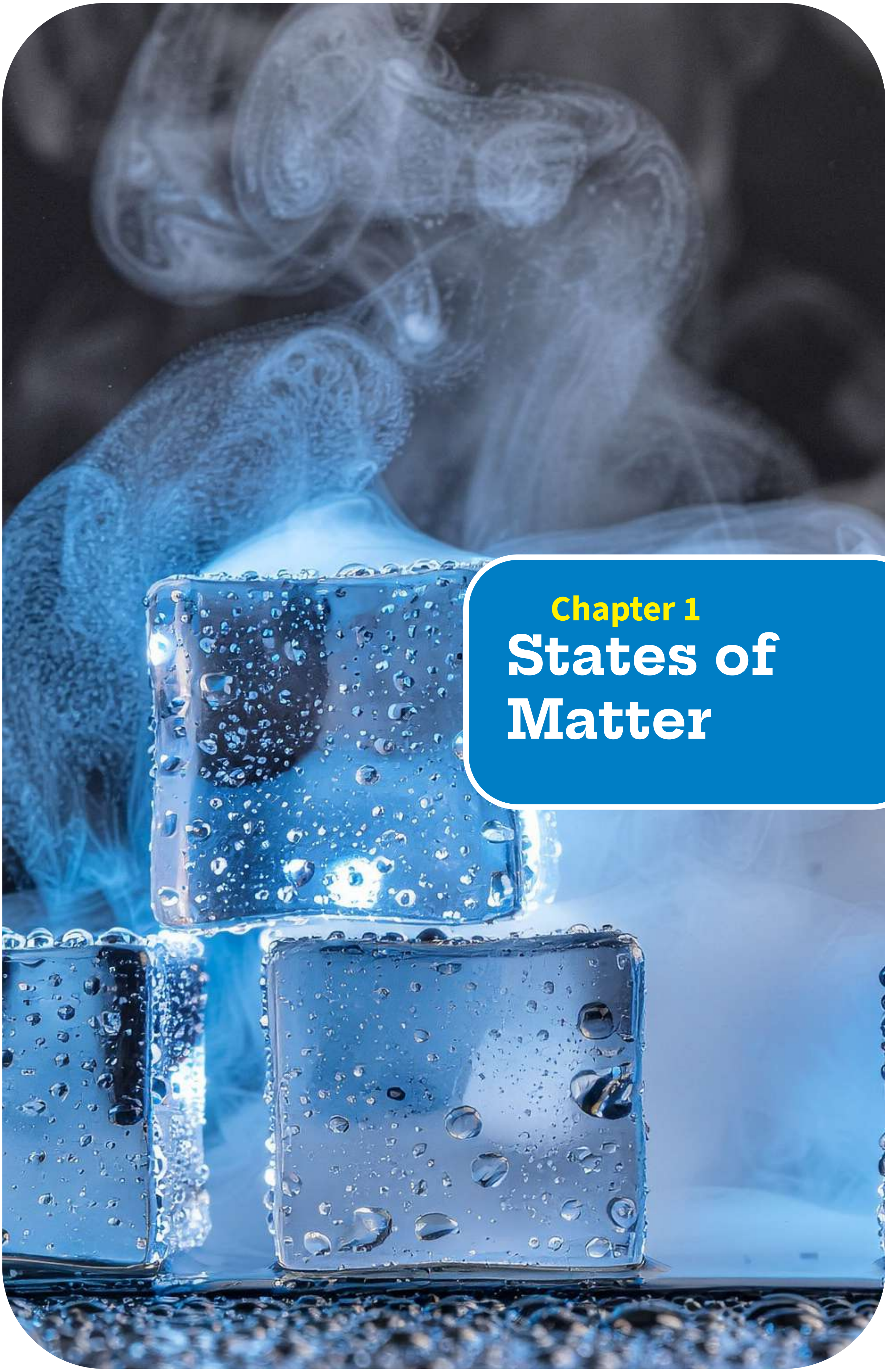
First Part — every lesson mapped to its EB Chemistry topic

Chapter 1 -States of Matter

Chapter 2 - Change and Equilibrium of Matter

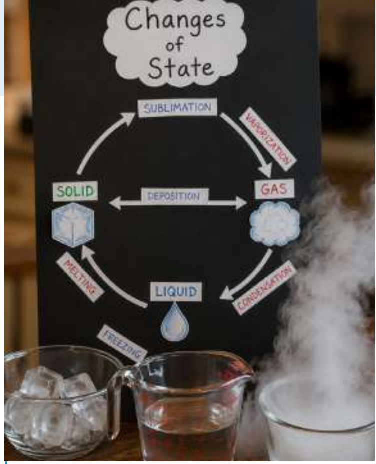
Chapter 1

States of Matter



1-1 — Changes of State

✨ Engaging Introduction



Every day, matter changes around us in fascinating ways. Ice melts into water, puddles disappear after rain, and clouds form high in the sky. Although these changes seem simple, they are driven by the movement of particles and the transfer of energy.

In this lesson, you will explore how and why substances change their state, investigate the forces acting between particles, and discover how these ideas help explain natural phenomena that shape our world.

❓ **Guiding question**: How does energy shape the states of matter around us?

📌 Point!

1. Celsius Temperature and Absolute Temperature

(a) **Celsius temperature:** the temperature scale used in everyday life, measured in degrees Celsius (symbol: °C).

(b) **Absolute temperature:** a temperature scale that starts at absolute zero (−273 °C) and uses the same interval size as the Celsius scale. The unit is the kelvin (symbol: K).

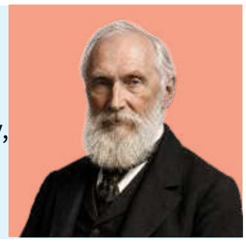
- The relationship between absolute temperature T (K) and Celsius temperature t (°C) is: $T = t + 273$.

Equation box

Relation	Expression	Note
Absolute ⇔ Celsius temperature	T (K) = t (°C) + 273	Same interval size; the zero point differs by 273

🏅 Appreciating Scientists

The Kelvin scale is named after Lord Kelvin, whose work helped scientists understand the relationship between temperature, energy, and particle motion. Today, the Kelvin scale is used worldwide in chemistry, physics, and space science.



2. Changes of State

(a) A substance can change between the solid, liquid, and gas states through melting, boiling, condensation, and freezing.

(b) Changes of state (see Figure 1.1.1):

(i) **Heat of fusion:** the amount of heat required to change 1 mol of a solid into a liquid.

(ii) **Heat of vaporization:** the amount of heat required to change 1 mol of a liquid into a gas.

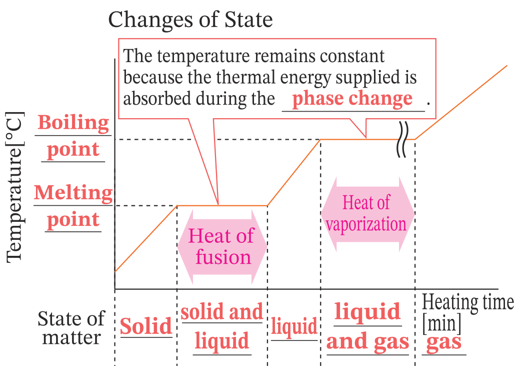


Figure 1.1.1. Relationship between temperature change and state of matter

(c) How to calculate the amount of heat

(i) Amount of heat in changes of state: Heat of fusion and heat of vaporization are expressed in kilojoules per mole [kJ/mol] measured at 1 atm pressure.

⭐ **Structure of Science**

Key Fact :
Changes of state and melting/boiling points depend on energy transfer and the strength of attractive forces between particles.

Key Concepts:

- Celsius temperature
- Absolute temperature
- Heat of fusion
- Heat of vaporization
- Van der Waals forces
- Hydrogen bond

Example — calculate the amount of heat required to melt *36 g* of ice at *0 °C*. (Given: heat of fusion of water = *6.0 kJ/mol; molar mass of H2O = 18 g/mol.)*
The Solution: 36 g of ice = 36 g ÷ 18 g/mol = 2.0 mol. Therefore the heat is *6.0 [kJ/mol] × 2.0 (mol) = 12 kJ*.

(ii) **Amount of heat in temperature changes:** Amount of heat in temperature changes: The specific heat capacity of water is 4.2 J/(g·°C) — that is, 4.2 J of heat is required to raise the temperature of 1 g of water by 1°C.

Example — The amount of heat required to heat 100 g of water from *5 °C to 25 °C is 4.2 [J/(g·°C)] × 100 (g) × [25 − 5] [°C] = 8400 J = 8.4 kJ.* $Q=m.\ c.\ (t_2-t_1)$

Think as an Engineer

You are designing a cooling system for transporting vaccines to remote communities. Which property would be more important when selecting a cooling substance?

☐ High heat of fusion ☐ High heat of vaporization

Give evidence on your answer

____

____

3. Forces between Particles and Melting / Boiling Points

(a) **Intermolecular forces:** attractive forces acting between molecules. They are **weaker** than covalent bonds, ionic bonds, and metallic bonds.

(i) **Van der Waals forces:** weak attractive forces between molecules. The greater the relative molecular mass, the **stronger** they are. For molecules of similar relative molecular mass, the forces between polar molecules are **stronger** than those between non-polar molecules.

(ii) **Hydrogen** bond: a bond formed between two highly electronegative atoms (F, O, N) via a hydrogen atom that is covalently bonded to one of them. Hydrogen bonds are stronger than Van der Waals forces.

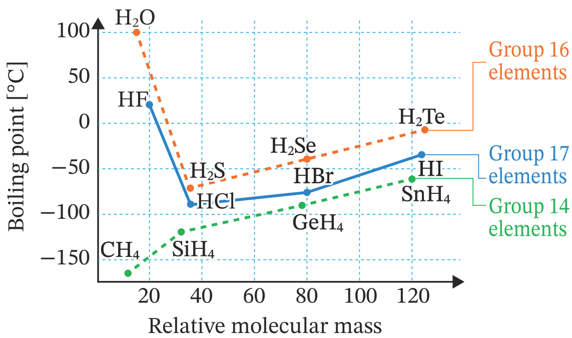


Fig 1.1.2. the boiling point curve

(b) **Forces between particles and melting / boiling points:** the stronger the attractive forces between the constituent particles, the higher the melting point and boiling point.

Example — For the boiling points of the substances shown in the figure 1.1.2 on the right:

Because **hydrogen** bonds form between the molecules, $H_2O > H_2S$. ·
Because it has a larger relative molecular mass, HBr > HCl. ·
Because the hydrides of Group 14 elements are non-polar molecules, they have low boiling points.

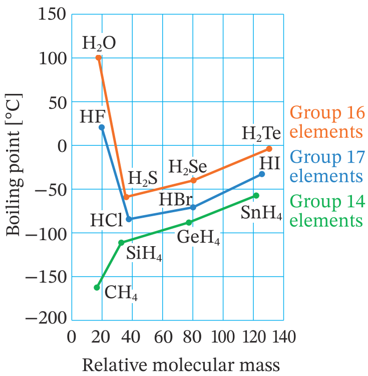
Warm Up

Answer the following questions:

(a) 90 g of water at 30 °C is heated until it turns completely into steam at 110 °C. How many kJ of heat are absorbed? Answer to two significant figures. Assume the heat needed to raise 1 g of water by 1 K is 4.2 J/(g·K), the heat needed to raise 1 g of steam by 1 K is 2.1 J/(g·K), and the heat of vaporization is 41 kJ/mol. (H = 1.0, O = 16)

____

(b) The figure below shows the boiling points of hydrides of Group 14, 16, and 17 elements. Referring to this figure, determine whether the underlined part of the following statement is True (T) or False (F). If it is False, correct the statement.



(i) Among hydrides of the same group in Periods 3 to 5, the <u>smaller</u> the relative molecular mass, the higher the boiling point.

(ii) Because the hydrides of Group 14 elements are <u>polar</u> molecules, the boiling points of the Group 14 hydrides are low within the same period.

(iii) Because hydrogen bonds form between HF molecules, the boiling point of HF is <u>higher</u> than that of HCl.

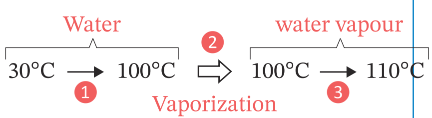
(c) Which substance has the higher boiling point, CO_2 or SiO_2? Answer with its chemical formula.

Be a Global Citizen

Access to clean water remains one of the greatest challenges facing humanity. Many countries — including Egypt, Saudi Arabia, Singapore, and Australia — use technologies based on evaporation and condensation to produce clean drinking water.

Explanation

(a) Noting that changes of state are occurring, calculate respectively:



❶ The amount of heat in the temperature change of water,

❷ The amount of heat in the change of state from water to steam,

❸ The amount of heat in the temperature change of steam.

> When expressing temperature differences, absolute temperature [K] and Celsius temperature [°C] yield the same numerical value.

❶ First, the amount of heat required to heat 90 g of water at 30 °C to 100 °C: 4.2 J/(g·K) × 90 g × (100 − 30) K = 26460 J = 26.46 kJ.

❷ Next, calculate the amount of heat required to turn the water into steam. Since the molar mass of water is 18 g/mol, 90 g of water is $\frac{90\,\text{g}}{18\,\text{g/mol}}$ = 5.0 mol.

> The heat involved in a change of state is considered in terms of the amount of substance [mol].

Therefore, the amount of heat required to turn 5.0 mol of water at 100°C into steam at 100°C is: 41 kJ/mol×5.0 mol = 205 kJ.

❸ Finally, the amount of heat required to heat 90 g of steam at 100°C to steam at 110°C is: 2.1 J/(g·K)×90 g×(110−100) K = 1890 J = 1.89 kJ.

Therefore, the total amount of heat absorbed is: 26.46 kJ+205 kJ+1.89 kJ = 233.35 kJ ≈ 2.3×10^2 kJ.

(b) (i) From the figure, it can be seen that for compounds of the same group in Periods 3 to 5, the larger the relative molecular mass, the higher the boiling point. → F: larger

(ii) The hydrides of Group 14 elements are non-polar molecules. Non-polar molecules have weaker Van der Waals

forces acting between them compared to polar molecules, so their boiling points are lower. → F: non-polar molecules.

(iii) **T**.

(c) CO_2 is a molecular crystal, and SiO_2 is a covalent network crystal (giant covalent structure). In covalent network crystals, ionic crystals, and metallic crystals, the forces acting between particles are stronger than intermolecular forces, so their melting points and boiling points are high. Therefore, SiO_2.

In molecules with similar structures, the greater the relative molecular mass the stronger the Van der Waals forces

Try

Choose the Correct Answer:

1. Two substances have similar relative molecular masses. Substance A boils at 100 °C, while Substance B boils at 35 °C. Which conclusion is most reasonable?

A. Substance A contains fewer molecules.

B. Substance A has stronger intermolecular forces.

C. Substance A has weaker covalent bonds.

D. Substance A contains larger atoms.

Answer: ____________

2. A puddle disappears several hours after rainfall. Which explanation best accounts for this observation?

A. The water molecules were destroyed.

B. The water changed into hydrogen and oxygen gases.

C. Water molecules gained energy to enter the gaseous state.

D. The water reacted with oxygen in the atmosphere.

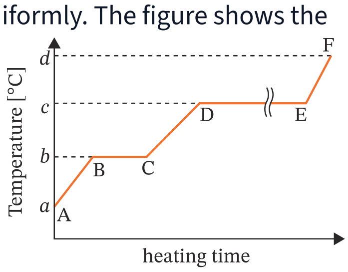
Answer: ____________

Solve the Following Questions :

3. How many kJ of heat are absorbed to turn 36 g of water at 50 °C into steam at 110 °C? Answer to two significant figures. Assume the heat to raise 1 g of liquid water by 1 °C is 4.2 J/(g·°C), the heat to raise 1 g of steam by 1 °C is 2.1 J/(g·°C), and the heat of vaporization of water at 100 °C is 41 kJ/mol. (H = 1.0, O = 16)

____

4. A solid of substance X at temperature a (°C) is heated uniformly. The figure shows the temperature change until it becomes a gas at temperature d (°C). Answer the following.



(a) What is the phenomenon during the constant-temperature interval BC called? Also state the state of substance X at that time.

____

(b) What are the temperatures b (°C) and c (°C) called, respectively?

____

(c) Heat is added during intervals BC and DE, yet the temperature does not rise. What is the heat being used for?

____

(d) What are the amounts of heat added during intervals BC and DE called, respectively?

____

Exercise

Choose the Correct Answer::

1. Which substance is expected to have the highest boiling point?

A. CH_4 B. H_2S

C. H_2O D. CO_2

Answer: ___________________________________

2. A scientist observes that Substance X melts at a much higher temperature than Substance Y. What can be inferred?

A. Substance X has stronger attractive forces.

B. Substance X contains fewer particles.

C. Substance X has a lower molecular mass.

D. Substance X contains more gas particles.

Answer: ___________________________________

Answer on the Following Questions:

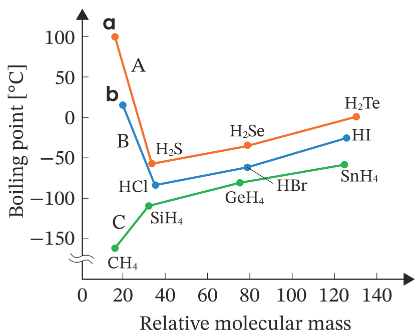
3. How many kJ of heat are required to heat 9.0 g of ice at 0 °C to water at 80 °C? Answer to two significant figures. Assume the heat of fusion of ice is 6.0 kJ/mol and the specific heat capacity of water is 4.2 J/(g·K). (H = 1.0, O = 16)

____

4. The figure shows the relationship between relative molecular mass and boiling point, and the groups A to C are hydrides of elements in the same group. Answer the following questions:

(a) Write the chemical formulas of the molecules corresponding to a and b.

____



(b) In molecules with very similar structures and bonding like those in group C, generally, the larger the relative molecular mass, the higher the melting point and boiling point. State the reason for this.

____

(c) Comparing the three compounds with the smallest relative molecular masses in groups A, B, and C, the boiling points of a and b are considerably higher compared to CH_4 in C. State the reason why such a difference occurs..

____

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

Challenge Your Thinking

Water has a much lower molecular mass than H_2Te, yet water has a much higher boiling point.

____

Can molecular mass alone explain boiling point?

____

Develop a scientific explanation for your answer

____

1-2 — Vapour-Liquid Equilibrium and Vapour Pressure

✨ Engaging Introduction

When we make tea, or boil water, bubbles suddenly appear and rise to the surface. Have you ever wondered what is inside these bubbles? Are they filled with air, water vapour, or something else? And why do they form only when the water reaches a certain temperature?

In this lesson, you will investigate the relationship between vapour pressure, boiling, and vapour–liquid equilibrium to uncover the science behind this everyday phenomenon.

> ❓ **Guiding question:** How do particle motion and intermolecular forces determine whether a liquid evaporates, condenses, or boils?

📌 **Point!**

1. Units of pressure

(a) 1 Pascal (Pa): The pressure when a force of 1 Newton (N) acts on an area of 1 m^2.
(b) 1 Hectopascal (hPa) = 100 Pa
(c) 1 Atmosphere (atm) = 1.013×10^5 Pa = 1013 hPa.

2. Vapour-Liquid Equilibrium and Vapour Pressure

A. Vapour-Liquid Equilibrium

When a liquid is placed in a closed container and left alone, two processes occur simultaneously: Evaporation, where molecules escape from the liquid surface, and Condensation, where vapour molecules return to the liquid surface.

· Kinetic energy of liquid molecules

The molecules composing a liquid do not all have the same kinetic energy.

Relationship between temperature and energy: As the temperature increases, the proportion of «high-energy molecules» able to overcome intermolecular forces and evaporate increases. Consequently the rate of evaporation (the number of molecules evaporating per unit time) increases.

· Vapour-liquid equilibrium as dynamic equilibrium

Inside a closed container the number of gaseous molecules increases over time, so the rate of condensation gradually rises. Eventually a state is reached where the **rate of evaporation equals the rate of condensation**.

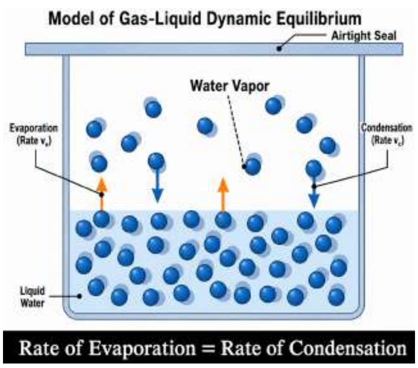


Fig 1.2.1 the relation between rate of evaporation and rate of condensation

⭐ **structure of science**

Key fact :

A liquid in a closed container reaches a state where evaporation and condensation occur at equal rates, called vapour–liquid equilibrium.

Key concepts:

- Vapour pressure
- Saturated vapour pressure
- Evaporation Condensation
- liquid–vapour equilibrium
- Boiling
- Atmospheric pressure
- Gas pressure
- Vacuum.

Although it appears as if both evaporation and condensation have stopped, molecular motion has not actually ceased. This state is called Vapour-Liquid Equilibrium.

B. Vapour Pressure

The pressure exerted by the gas at vapour-liquid equilibrium is called the Saturated Vapour Pressure, or simply the Vapour Pressure.

Characteristics of vapour pressure

- **Temperature dependence:**

  As the temperature increases, the number of molecules with high kinetic energy increases, and thus the vapour pressure increases. A graph illustrating this relationship is called the Vapour Pressure Curve.

- **Substance dependence:**

  substances with weaker intermolecular forces evaporate more easily and so have a higher vapour pressure at the same temperature (e.g., Diethyl ether > Ethanol > Water).

- **Independence from volume:** The vapour pressure remains constant as long as the temperature is kept constant and liquid is present, regardless of changes in the amount of liquid or the volume of the gas in the container.

C. Boiling

A phenomenon where gas is actively generated not only from the surface of the liquid but also from its interior.

Boiling occurs when the vapour pressure equals the external pressure (atmospheric pressure).

> **Example —** from the graph, at an atmospheric pressure of 1.013×10^5 Pa water boils at 100 °C.

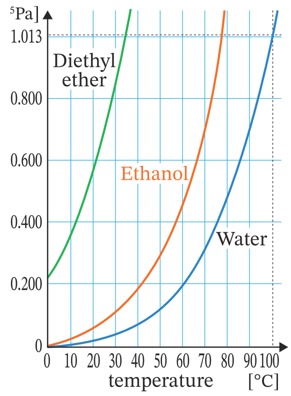


Fig 1.2.2 the relation between vapour pressure and temperature

Think Critically — Claim · Evidence · Reasoning (CER)

Build your argument.

Claim: a liquid with a higher vapour pressure evaporates more readily.
Evidence: use information from the vapour-pressure graph.
Reasoning: explain how particle motion and intermolecular forces support this claim.

____

Life Application Pressure cookers: water boils when its vapour pressure equals the surrounding atmospheric pressure. By using a sealed lid, a pressure cooker traps steam, raising the pressure and therefore the boiling point of water — so food cooks significantly faster at higher temperatures.

3. Mercury column and gas pressure

(a) **Mercury column:** Mercury column: When a glass tube closed at one end is filled with mercury and inverted into a container containing mercury, the mercury column stabilises at a certain height as shown in the right figure 1.2.3, and a mercury column is formed in the tube. At this time, the upper part becomes a vacuum.

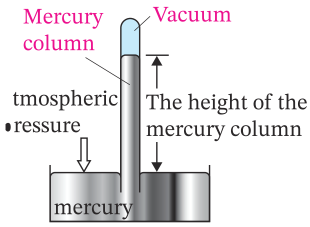


Fig 1.2.3 measuring atmospheric presure

(b) **Mercury column and pressure**

- The height of the mercury column is directly proportional to the atmospheric pressure.
- The height of the mercury column is 760 mm at 1.013×10^5 Pa (1 atm).

⇒ Using the millimetre of mercury (symbol: mmHg),

760 mmHg = 1.013×10^5 Pa.

- If gas is introduced into the upper part of the column, the height decreases; the amount of decrease is directly proportional to the pressure of the gas.

Pressure units at a glance

Unit	Equivalent
1 atm	1.013×10^5 Pa = 1013 hPa = 760 mmHg
1 hPa	100 Pa
Boiling condition	vapour pressure = external (atmospheric) pressure

Experimental Investigation

Build Your Own Atmospheric Pressure Indicator

Materials

· Empty glass
· jar or plastic container
· balloon (or biodegradable rubber sheet)
· rubber band
· drinking straw (paper preferred
· cardboard or recycled paper
· glue or tape
· marker.

Safety precautions

Use scissors carefully when cutting
· ensure the glass jar is free from cracks
· keep the device away from excessive heat
· work under teacher supervision when using sharp tools.

Procedure

1. Cut the balloon and stretch it tightly over the opening of the jar.
2. Use a rubber band to hold it firmly in place.
3. Tape one end of the straw to the centre of the stretched balloon surface.
4. Place a piece of cardboard behind the straw and draw a simple scale with reference marks.
5. Keep the device in the same location and record the straw›s position over several days.

How it works

Changes in atmospheric pressure push on the flexible balloon membrane.
• Higher atmospheric pressure pushes the membrane downward.
• Lower atmospheric pressure allows the membrane to move upward.
As the membrane moves, the straw acts as a pointer, indicating changes in atmospheric pressure.

Warm Up

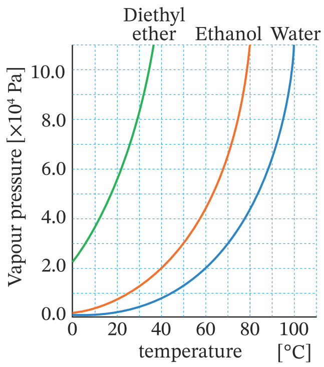
Answer the following questions.

(a) The graph represents the vapour-pressure curves of diethyl ether, ethanol, and water. Answer the following questions regarding these:

(i) What is the boiling point of ethanol when the atmospheric pressure is 1.0×10^5 Pa?

(ii) Among the three substances, name the one with the weakest intermolecular forces.

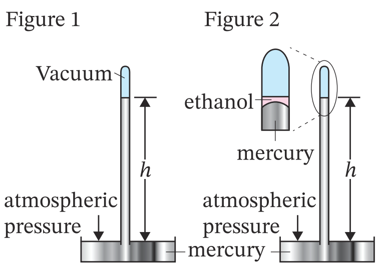
(iii) Among the three substances, which is least likely to evaporate at 20 °C? Name the substance.



(b) Under 25°C and 1.013×10^5 Pa, when a glass tube of 1 m length closed at one end is filled with mercury and set as shown in Figure 1, the height h of the mercury in the glass tube becomes 760 mm, and the upper part of the mercury column becomes a vacuum. The atmospheric pressure at this time is expressed as 760 mmHg. Answer the following questions:

(i) When the atmospheric pressure is 3.039×10^4 Pa, what will h be in mm?

(ii) As shown in Figure 2, when a small amount of ethanol is introduced into the vacuum part of the glass tube and it reaches a state of vapour-liquid equilibrium, what will h be in mm at an atmospheric pressure of 1.013×10^5 Pa? Assume the saturated vapour pressure of ethanol at 25°C is 65 mmHg, and the mass of ethanol can be ignored.



Explanation

(a) (i) Since the temperature at which the atmospheric pressure and the vapour pressure become equal is the boiling point, just read the temperature when the vapour pressure is 1.0×10^5 Pa. From the graph, the temperature of ethanol when the vapour pressure is 1.0×10^5 Pa = 10.0×10^4 Pa is 78°C.

(ii) The weaker the intermolecular forces, the lower the boiling point. The substance with the lowest boiling point under 10.0×10^4 Pa is diethyl ether. Therefore, diethyl ether.

(iii) The greater the vapour pressure, the more molecules in the liquid escape to the outside, and evaporation occurs.

Therefore, the least likely to evaporate is water, which has the lowest vapour pressure at 20°C according to the graph.

(b) (i) The height of the mercury column is directly proportional to the atmospheric pressure. When the atmospheric pressure is 1.013×10^5 Pa, h is 760 mm, so h when the atmospheric pressure is 3.039×10^4 Pa is:

1.013×10^5 Pa : 760 mm = 3.039×10^4 Pa : h [mm].

h = 228 mm: Answer 228 mm

(ii) When ethanol is introduced into the upper part and it reaches a state of vapour-liquid equilibrium, inside the

Chemistry Around the World

The way people cook can be shaped by their environment. In high-altitude regions, lower atmospheric pressure causes water to boil below 100 °C. To overcome this, many households use pressure cookers, which raise the pressure inside the container and let food cook more efficiently.

tube, the saturated vapour pressure of ethanol pushes the mercury down by 65 mmHg. Since h = 760 mm when

the upper part is a vacuum at an atmospheric pressure of 1.013×10^5 Pa, h when ethanol is introduced is:

h = 760 mm−65 mm Answer : 695 mm.

Try

Solve the Following Questions:

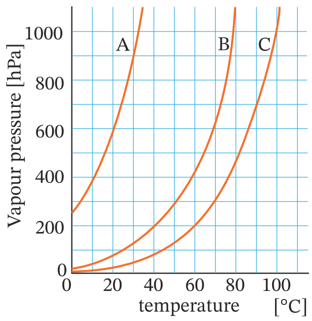
1. The figure shows the vapour-pressure curves of substances A to C. Answer the following.

(a) Which of A to C has the highest boiling point?

(b) Which of A to C has the strongest intermolecular forces?

(c) At what temperature does substance B boil when the external pressure is 800 hPa?

(d) Which of A to C is water?



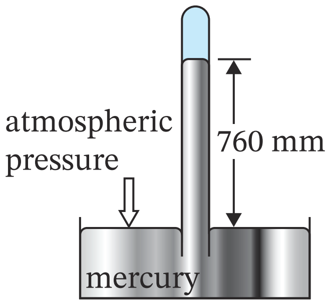
Answer: ____________________________________

2. When a glass tube of 1 m length closed at the upper end is filled with mercury and inverted into a mercury trough as shown in the figure at an atmospheric pressure of 1.0×105 Pa, the height of the mercury column in the tube was 760 mm. Answer the following questions with two significant figures.

(a) State the condition of the space in the upper part of the tube.

(b) After a cold front, the atmospheric pressure drops and the column becomes 745 mm. What is the atmospheric pressure in Pa?

(c) A small amount of gas is introduced from the bottom and the column drops to 494 mm. What is the gas pressure in the tube in Pa?



Answer: ____________________________________

Exercise

Choose the correct answer:

1. Why is atmospheric pressure lower at higher altitudes?

A. Air molecules become larger.

B. The atmosphere contains less oxygen.

C. There are fewer air molecules above the location.

D. Water vapour disappears.

Answer: ____________________________________

Answer on the Following Questions

2. The figure shows the vapour-pressure curves of substances A to C. Answer the following.

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

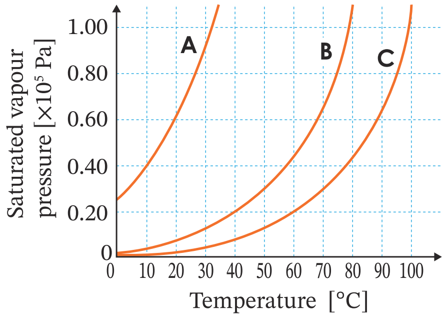
....

(a) From the figure, find the boiling point of A at 1.0×10^5 Pa.

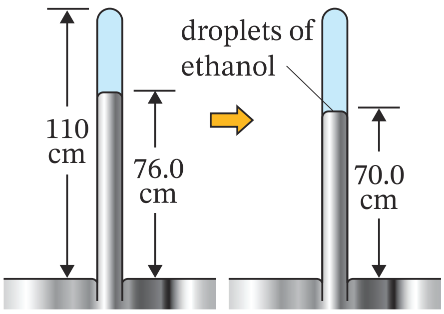
(b) What is the boiling point of water at the summit of Mount Sinai (atmospheric pressure 0.75×10^5 Pa)?

Answer: ___________________________________

(c) Write the letter of the substance with the weakest intermolecular forces.



3. Under 1.01×10^5 Pa and 25 °C, a glass tube closed at one end is filled with mercury and inverted into a trough, with 110 cm of tube above the mercury surface; the mercury column settles at 76.0 cm. When ethanol is injected it all evaporates and the column falls; with more ethanol it falls further until, when liquid ethanol droplets remain, the column becomes constant at 70.0 cm. What is the saturated vapour pressure of ethanol at 25 °C in Pa? Answer with two significant figures.



Answer: ___________________________________

1-3 — Phase Diagram

✨ Engaging Introduction

Water can exist as ice, liquid water, or water vapour. Under certain conditions all three states can exist at the same time, and at extremely high temperatures and pressures the distinction between liquid and gas disappears completely. So how can a single substance behave in such different ways?

In this lesson, you will learn how chemists use phase diagrams to predict the state of matter and explain some of the most unusual properties of substances.

❓ **Guiding question**: How can a phase diagram help us predict and explain the physical state of a substance under different conditions?

Point!

Key concept: the phase diagram of a pure substance

The state (solid, liquid, or gas) of a pure substance is determined by the conditions of temperature and pressure. A phase diagram is a graphical representation that shows the state a substance exists in at a given temperature and pressure.

The phase diagram literacy

Study the phase diagram and answer:

• Which curve separates liquid and gas?

Answer: ___________________________________

• Which curve separates solid and liquid?

Answer: ___________________________________

• Which curve separates solid and gas?

Answer: ___________________________________

Support your answers using evidence from the diagram.

Answer: ___________________________________

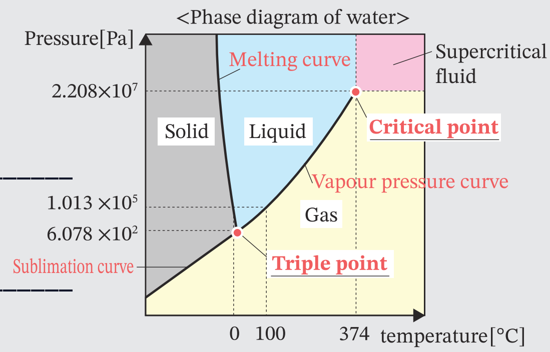


Fig 1.3.1 Phase diagram of Water

1. The three regions and three curves of a phase diagram

A phase diagram is divided into three regions by three curves, with each region corresponding to one of the three states: solid, liquid, or gas. Along the curves that separate these states, the two adjacent states coexist in equilibrium at that specific temperature and pressure.

• Melting Curve(or Fusion Curve): This is the boundary between the solid and liquid states.

• Vapour Pressure Curve: This is the boundary between the liquid and gaseous/vapour states.

• Sublimation Curve : This is the boundary between the solid and gaseous/vapour states.

⭐ **Structure of Science**

Key Fact :

A phase diagram shows the physical state (solid, liquid, or gas) of a pure substance at different temperatures and pressures.

Key Concepts:

the phase diagram of a pure substance

Predict the State

A substance is located at a point inside the liquid region. Predict what will happen if:

• Temperature increases while pressure remains constant.

____

• Pressure decreases while temperature remains constant.

____

Use the phase diagram to justify your answer

____

2. The triple point and the critical point

Important "points" exist on the phase diagram that demonstrate the unique properties of a substance.

- Triple Point: This is the point where the three curves intersect. At this specific temperature and pressure, the three states-solid, liquid, and gas-coexist simultaneously.

- Critical Point and Supercritical Fluid: The critical point is the specific temperature and pressure at the upper end of the vapour pressure curve, beyond which the liquid and gas phases become indistinguishable.

At temperatures and pressures higher than this point (supercritical conditions), the substance exists in a state ,where the distinction between liquid and gas ceases, exhibiting properties intermediate between the two. A substance in this state is called supercritical fluid.

Reading a phase diagram

Feature	What it marks
Melting curve	Solid ⇔ liquid boundary
Vapour-pressure curve	Liquid ⇔ gas boundary
Sublimation curve	Solid ⇔ gas boundary
Triple point	Solid, liquid, and gas coexist
Critical point	Liquid and gas become indistinguishable

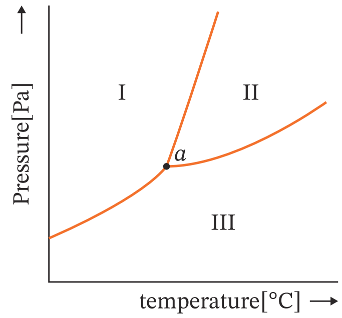
Life Application

Supercritical Fluids in Daily Life and Industry Supercritical fluids have a remarkable property: they can penetrate spaces like a gas while dissolving substances like a liquid. For instance, supercritical carbon dioxide (CO_2) is used to extract only caffeine from coffee beans (decaffeination). This technology is also used worldwide to extract perfume ingredients from Egyptian-grown jasmine and roses without damaging their delicate, heat-sensitive aromas.

Warm Up

The figure schematically shows the phase diagram of carbon dioxide. Answer the following questions.

(a) In regions I, II, and III, in what state is carbon dioxide respectively? Choose the correct combination from **A** to **D**.



A. I: Gas II: Liquid III: Solid

B. I: Solid II: Gas III: Liquid

C. I: Solid II: Liquid III: Gas

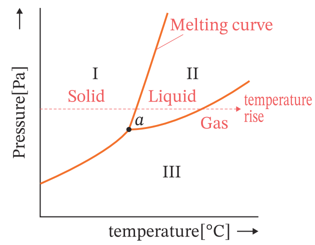
D. I: Liquid II: Gas III: Solid

(b) Write the name of the change of state from region II to region I.

(c) How does the melting point of solid carbon dioxide change as the pressure increases?

(d) What is a point a called?

Explanation



(a) Consider that as the temperature increases at a constant pressure, the state changes from solid to liquid to gas.
Answer: C

(b) It is a change of state from liquid to solid.
Freezing (Solidification).

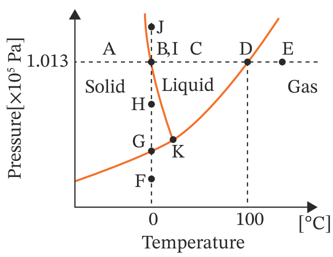
(c) The temperature at which a solid becomes a liquid is the melting point. From the melting curve in the figure, it can be seen that as the pressure increases, the melting point becomes higher.

(d) triple point .

Try

Solve the Following Questions:

1. Fill in the blanks (I) to (VIII) with the appropriate scientific terms, the phase diagram of water provided.



a) When heated from A to E, the solid begins to turn into a liquid by (I) at 0 °C.

The (II) remains constant until all the solid becomes liquid, but after all the

solid has undergone (I), the temperature rises again and the liquid state

continues up to D. The temperature at D is called (III), and that temperature is 100°C. If heating continues, (II) remains constant until all the liquid has (IV) to from (V).

b) As pressure is increased from F to J, at G, the gas begins to form a solid by (VI), at H, it all becomes solid, and at I, the solid (VII) to form a liquid. In this case, the density of the liquid at point J is greater than that of the solid at point H. Furthermore, the point K where the curves intersect is called (VIII).

Grand Challenge

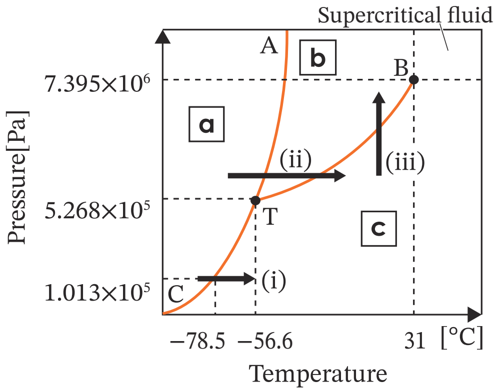
Designing a Martian Habitat Future astronauts will need reliable sources of water. Using your understanding of phase diagrams, design a system that keeps water in the liquid state on Mars. Consider: temperature · pressure · engineering solutions. Present your design and a scientific justification.

Exercise

1. The figure is the phase diagram of carbon dioxide. Based on it, answer the following Questions.

(a) In what state are regions a to c, respectively?

Answer: ___________________________________



(b) What are the curves AT, BT, and CT called, respectively?

Answer: ___________________________

(c) What is point T, where the three curves intersect, called?

Answer: ___________________________

(d) What is the change of state indicated by arrow (i) called?

Answer: ___________________________

(e) Briefly explain the change of state that occurs when the temperature is raised as in arrow (ii).

Answer: ___________________________________

(f) What is the change of state that occurs when the pressure is raised as in arrow (iii) called?

Answer: ___________________________________

(g) At temperatures and pressures higher than point B, the states become indistinguishable. What is point B called?

Answer: ___________________________________

Challenge Your Thinking

Draw the Journey A substance starts in the solid region.

Draw a pathway on the phase diagram that would cause it to:

1. Melt 2. Boil 3. Condense 4. Freeze

Explain each change on your diagram

____

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

1-4 — Gas Laws

Engaging Introduction

On a hot summer day, a bicycle tyre may feel firmer than it did on a cool morning, even though no air has been added or removed. How could a change in temperature affect the behaviour of the gas inside the tyre?

In this lesson, you will explore how pressure, volume, and temperature are related, and how chemists use mathematical models to predict the behaviour of gases in everyday life and in technology.

❓ Guiding question: How can mathematical relationships help us explain and predict the behaviour of gases under changing conditions?

Exploration

Consider:

• A bicycle tire left in the sun.

• A weather balloon rising through the atmosphere.

• A sealed gas container placed in a freezer.

Without using equations:

• What variables might be changing?

____

• What effects do you predict?

____

Point!

Gas Laws

(a) Boyle's law:

at constant temperature, the volume V of a given amount of gas is inversely proportional to its pressure P.

At a constant temperature, when a gas with pressure P_1 and volume V_1 changes to pressure P_2 and volume V_2, the following equation applies ● $P_1V_1 = P_2V_2$. ● PV = constant.

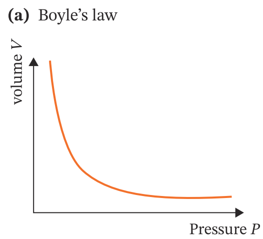


Fig 1.4.1 Boyle`s Law

(b) Charles's law:

At a constant pressure, the volume V of a given amount of gas is directly proportional to its absolute temperature T.

• At a constant pressure, when a gas with absolute temperature T_1 and volume V_1 changes to absolute temperature T_2 and volume V_2, the following equation applies: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ ____________ $\frac{V}{T}$

(c) Combined gas law (Boyle–Charles's law):

The volume V of a given amount of gas is inversely proportional to its pressure P and directly

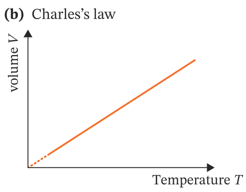


Figure 1.4.2 Charles`s law

Structure of Science

Key principles

At constant temperature, the volume of a gas is inversely proportional to its pressure (Boyle's law). - At constant pressure, the volume of a gas is directly proportional to its absolute temperature (Charles's law).

Misconception Alert

Wrong

«If pressure increases, gas particles become smaller.»

Right

Gas particles do not shrink. Instead, the space between particles decreases.

proportional to its absolute temperature T.

- When a given amount of gas with volume V1, pressure P1, and absolute temperature T1 changes to volume V2 , pressure P2, and absolute temperature T2, the following equation applies: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$.

(d) **Application of the gas laws**

- To find the volume, pressure, or temperature of a given amount of gas, **the combined gas law** is used

- When using the gas laws, the temperature must be expressed as **absolute temperature [K]**.

Equation box

Law	Condition	Relation
Boyle's law	constant T	$P_1V_1 = P_2V_2$ (PV = constant)
Charles's law	constant P	$V_1 / T_1 = V_2 / T_2$ (V / T = constant)
Combined gas law	—	$P_1V_1 / T_1 = P_2V_2 / T_2$ (PV / T = constant)

Appreciating Scientists

In the seventeenth century, Robert Boyle carefully investigated the relationship between pressure and volume. His experiments demonstrated that scientific laws can emerge from systematic observations, precise measurements, and mathematical analysis.

Think Like a Chemist

Chemists rarely measure every property of a gas directly. Instead, they use established relationships to predict unknown quantities from known measurements.

When solving a gas-law problem, ask yourself:

- Which variables are changing?

____

- Which variables remain constant?

____

- Which gas law best describes the situation?

____

Scientific Application: Designing Safe Gas Container

Engineers must consider changes in pressure and temperature when designing gas cylinders used in hospitals, laboratories, and industries.

- An unsafe design could lead to dangerous pressure build-up.
- Gas laws help in predicting these changes and design safer technologies.

Warm Up

Answer the following questions.

(a) A sample of gas has a volume of 200 mL at a pressure of 4.0×10^4 Pa. If the volume is changed to 1.0 L at constant temperature, what is the new pressure in Pa? (Two significant figures.)

(b) A sample of gas occupies 60 L at 27 °C and 1.0×10^5 Pa. If the temperature is changed to 87 °C and the pressure to 2.5×10^5 Pa, what will the new volume be in L? (Two significant figures.)

Explanation

(a) Since we need to find the pressure of a given amount of gas, we use combined gas law:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}.$$

Let the required pressure be P [Pa].

The pressure before the change is P1 = 4.0×10^4 Pa, and the volume is V_1 = 200 mL = 0.200 L.

The pressure after the change is P_2 = P [Pa], and V_2 = 1.0 L. Therefore,

$$\frac{4.0 \times 10^4\text{Pa} \times 0.200\text{L}}{T_1} = \frac{P[\text{Pa}] \times 1.0\text{L}}{T_2}$$

Also, since the temperature is constant, $T_1 = T_2$, so 4.0×10^4 Pa×0.200 L = P [Pa]×1.0 L.

$$P = \frac{4.0 \times 10^4\text{Pa} \times 0.200\text{L}}{1.0\text{L}}$$

= 8.0×10^3 Pa

(b) Since we need to find the volume of a given amount of gas, we use combined gas law: $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

Let the required volume be V [L].

The temperature before the change is T_1 = (27+273) K, the pressure is P_1 = 1.0×105 Pa, and the volume is V_1 = 60 L.

The temperature after the change is T_2 = (87+273) K, the pressure is P_2 = 2.5×105 Pa, and the volume is V_2 = V [L]. Therefore,

$$\frac{1.0 \times 10^5\text{Pa} \times 60\text{L}}{(27+273)\text{K}} = \frac{2.5 \times 10^5\text{Pa} \times V[\text{L}]}{(87+273)\text{K}}$$

$$\frac{1.0 \times 10^5\text{Pa} \times 60\text{L}}{300\text{K}} = \frac{2.5 \times 10^5\text{Pa} \times V[\text{L}]}{360\text{K}}$$

$$V = \frac{1.0 \times 10^5\text{Pa} \times 60\text{L} \times 360\text{K}}{300\text{K} \times 2.5 \times 10^5\text{Pa}}$$

$$= \frac{72}{2.5}$$

= 28.8

≈ 29 L

NOS: Why Do Scientific Models Matter?

Gas laws do not describe individual gas particles. Instead, they provide models that help scientists explain and predict the collective behavior of millions of particles.

Try

Answer the following questions:

(a) When 4.0 L of a gas at 1.0×10^5 Pa is subjected to a pressure of 2.0×10^5 Pa at constant temperature, what will the volume be in L? (Two significant figures.)

Answer: ______________________________

(b) To expand 2.00 L of a gas at 27 °C and 2.0×10^5 Pa to 2.50 L at constant pressure, what should the temperature be in °C? (Nearest whole number.)

Answer: ______________________________

(c) When a gas at 0 °C and 1.0×10^5 Pa is raised to 97 °C at constant volume, what will the pressure be in Pa? (Two significant figures.)

Answer: ______________________________

(d) When 500 mL of a gas at 77 °C and 3.0×10^5 Pa is changed to 47 °C and 2.0 L, what will the pressure be in Pa? (Two significant figures.)

Answer: ______________________________

Exercise

Choose the correct answer:

1. A balloon is cooled while the pressure remains constant. According to Charles's law:

A. Volume increases. B. Volume decreases.

C. Pressure increases. D. Number of moles increases.

Answer: ______________________________

2. A student calculates a gas volume using the combined gas law and obtains a value larger than the initial volume. Which situation could explain this result?

A. Pressure increased while temperature stayed constant.

B. Temperature increased while pressure decreased.

C. Temperature decreased while pressure increased.

D. Both temperature and volume decreased.

Answer: ______________________________

3. Answer the following questions:

(a) When 5.0 L of a gas at 1.0×10^5 Pa is compressed to 1.0 L at constant temperature, what will the pressure be in Pa? (Two significant figures.)

(b) When 320 mL of a gas at 10 °C and 5.0×10^5 Pa is subjected to 4.0×10^5 Pa at constant temperature, what will the volume be in mL? (Two significant figures.)

(c) When 3.0 L of a gas at 7 °C and 1.0×10^5 Pa is raised to 127 °C at constant pressure,

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

what will the volume be in L? (Two significant figures.)

(d) To expand 2.00 L of a gas at 7 °C to 6.00 L at constant pressure, what should the temperature be in °C? (Nearest whole number.)

(e) When 3.0 L of a gas at 7 °C and 1.5×10^5 Pa is raised to 127 °C at constant volume, what will the pressure be in Pa? (Two significant figures.)

(f) To raise the pressure of a gas at 27 °C and 2.00×10^5 Pa to 2.20×10^5 Pa at constant volume, what should the temperature be in °C? (Nearest whole number.)

(g) When 2.0 L of a gas at 23 °C and 1.5×10^5 Pa is changed to 117 °C and 3.0 L, what will the pressure be in Pa? (Two significant figures.)

(h) When 600 mL of a gas at 17 °C and 2.0×10^5 Pa is changed to 0 °C and 9.1×10^4 Pa, what will the volume be in L? (Two significant figures.)

Answer: ________________________________

Challenge Your Thinking

A weather balloon is launched from sea level and rises to an altitude where the atmospheric pressure is much lower than at the surface. Without performing calculations:

- Predict how the balloon's volume will change.

____

- Explain your reasoning using particle theory.

____

- Identify which gas law best supports your prediction.

____

-Sketch a simple diagram to communicate your thinking.

____

1-5 — Ideal Gas Equation

Engaging Introduction

A chemist discovers a sealed container filled with an unknown gas. By measuring only its pressure, volume, temperature, and mass, the chemist can identify important characteristics of the gas without ever observing its particles directly. Let's explore how chemists use measurements and mathematical models to uncover the secrets of gases.

❓ **Guiding question**: How can one equation help scientists explain and predict the behaviour of gases under different conditions?

Exploration

Consider the following situations:

- Bread dough rising in an oven.
- A sealed juice box expands when left in a hot car.
- A camping gas canister used in cold weather.
- A scuba diver breathing compressed air underwater.

Without using calculations:

- Which variables might change?
- Which variables might remain constant?
- What observations would you expect?

Point!

The Ideal Gas Law

The relationship between the state variables of a gas-pressure (P), volume (V), absolute temperature (T), and the amount of substance (n)-is consolidated into a single mathematical expression known as the **Ideal Gas Law** . This can be logically derived from the Combined Gas Law (Boyle-Charles' Law).

Step 1 — Starting with one mole of gas

First, recall the Combined Gas Law (Boyle-Charles' Law). For a fixed amount of gas, this law states that the product of its pressure and volume is directly proportional to its absolute temperature.

If we consider 1 mole of gas, we define the molar volume as v [L/mol] and the constant as R. This leads to the following equation:

$Pv = RT$

Step 2 — Determining the molar gas constant (R)

Irrespective of the gas species, 1 mole of any ideal gas occupies a molar volume of 22.4 L/mol at standard conditions (0 °C = 273 K, $1.01^3\times10^5$ Pa).

By substituting these standard values into the previous equation (Pv = RT), we can calculate the constant R, which is termed the Molar Gas Constant.

$$R = \frac{PV}{T} = \frac{1.013 \times 10^5\text{Pa} \times 22.4\text{L/mol}}{273\text{K}} = 8.31 \times 10^3 \frac{\text{Pa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$$

Structure of Science

Key Fact :

The Ideal Gas Law combines the patterns discovered by earlier scientists into a single equation that explains and predicts gas behaviour over a wide range of conditions.

Key Concepts:

· Pressure

· Volume

· Molar volume

· Molar mass

· Molar gas constant (R).

Step 3 — Extending to n moles (completing the Ideal Gas Law)

When the amount of substance is n [mol], the total volume V is n times the molar volume v (V = nv), so v = V ÷ n. Substituting into Step 1, then multiplying both sides by n and rearranging gives the complete equation of state: $PV = nRT$.

Application of the Ideal Gas Law: determination of molar mass

The Ideal Gas Law is also a key tool for finding the molar mass of an unknown gas. If ω [g] is the mass of the gas and M [g/mol] its molar mass (numerically equal to the relative molecular mass), then $n = \frac{\omega}{M}$. Substituting into the Ideal Gas Law gives the following relationship.

$$PV = \frac{\omega}{M}RT$$

Using this equation, you can calculate the molar mass (M) of the gas from experimentally measured values, such as its mass and volume.

Equation box

Quantity	Relation
Ideal gas law (1 mol)	Pv = RT
Molar gas constant	R = 8.31 × 10^3 Pa·L/(mol·K)
Ideal gas law (n mol)	PV = nRT
Molar mass form	PV = (ω / M) RT

Life Application

Heat and Car Tyres : Heatwaves and Car Tyres Take care to prevent car-tyre blowouts in summer, when the asphalt surface is extremely hot. The Ideal Gas Law, PV = nRT, explains the danger: with the amount of air (n) and the volume (V) inside the tyre roughly constant, a rapid rise in air temperature (T) from the hot road raises the internal pressure (P) proportionally,

Mathematics in Chemistry

Application of the Ideal Gas Law:

determination of molar mass Scientific equations are not simply mathematical shortcuts. They are models built from experimental evidence.

The Ideal Gas Equation allows scientists to:

- Predict behavior.
- Test hypotheses.
- Analyze data.
- Communicate results quantitatively.

Warm Up

Answer with two significant figures. Assume R = 8.3 × 10^3 Pa·L/(K·mol).

(a) Answer the following.

(i) There is 8.3 L of gas at 27 °C and 3.0 × 10^5 Pa. What is the amount of substance in mol?

(ii) When 1.0 mol of gas is expanded to 24.07 L at 17 °C, what is the pressure in Pa?

(b) A certain gas had a density of 0.53 g/L at 27 °C and 3.0 × 10^4 Pa. What is the relative molecular mass of this gas?

Explanation

(a) (i) To determine the amount of substance, we use the ideal gas equation. Let the required amount of substance be n [mol]. From the ideal gas equation PV = nRT:

3.0×10^5 Pa × 8.3 L = n [mol] × 8.3×10^3 Pa.L/(K.mol) × (27+273) K

● Convert the unit of temperature to [K].

$$n = \frac{3.0\times10^5\ \text{Pa}\times 8.3\ \text{L}}{8.3\times10^3\ \text{Pa·L/(K·mol)}\times(27+273)\ \text{K}}$$

$$= \frac{\overset{1}{\cancel{3.0}}\times\overset{1}{\cancel{10^5}}\ \text{Pa}\times \overset{1}{\cancel{8.3}}\ \text{L}}{\underset{1}{\cancel{8.3}}\times\underset{1}{\cancel{10^3}}\ \text{Pa·L/(K·mol)}\times \underset{1}{\cancel{300}}\ \text{K}}$$

= 1.0 mol

(ii) Since the amount of substance is given, we use the ideal gas equation. Let the required pressure be P [Pa]. From the ideal gas equation PV = nRT:

$$P\ [\text{Pa}]\times 24.07\ \text{L} = 1.0\ \text{mol}\times 8.3\times10^3\ \text{Pa·L/(K·mol)}\times(17+273)\ \text{K}$$

$$P = \frac{1.0\ \text{mol}\times 8.3\times10^3\ \text{Pa·L/(K·mol)}\times(17+273)\ \text{K}}{24.07\ \text{L}}$$

$$= \frac{1.0\ \text{mol}\times 8.3\times10^3\ \text{Pa·L/(K·mol)}\times 290\ \text{K}}{24.07\ \text{L}}$$

$$= \frac{\overset{1}{\cancel{24.07}}\times10^5}{\underset{1}{\cancel{24.07}}}$$

= 1.0×10^5 Pa

(b) The relative molecular mass is determined from the molar mass [g/mol].

Let the molar mass of this gas be M [g/mol]. Since the density is 0.53 g/L, the mass is 0.53 g when the gas volume is 1.0 L.

Therefore, the amount of substance of this gas is $\frac{0.53}{M}$ [mol].

Amount of substance [mol] = $\frac{\text{Mass [g]}}{\text{Molar mass [g/mol]}}$

Now that the amount of substance is known, using the ideal gas equation PV = nRT:

3.0×10^4 Pa × 1.0 L = $\frac{0.53}{M}$ [mol] × 8.3×10^3 Pa.L/(K.mol) × (27+273) K

$$M = \frac{0.53\times 8.3\times10^3\ \text{Pa·L/(K·mol)}\times(27+273)\ \text{K}}{3.0\times10^4\ \text{Pa}\times 1.0\ \text{L}}$$

$$= \frac{0.53\times 8.3\times\cancel{10^3}\ \text{Pa·L/(K·mol)}\times \overset{100}{\cancel{300}}\ \text{K}}{\cancel{3.0}\times\cancel{10^4}\ \text{Pa}\times 1.0\ \text{L}}$$

$= 0.53\times 8.3\times10$

$= 43.99$

≈ 44 g/mol

Therefore, the relative molecular mass of this gas is 44 g/mol

Try

Solve the following Questions:

1. Assume the gas constant R = 8.3×10^3 Pa·L/(K·mol).(O = 16)

(a) When 0.40 mol of gas is subjected to a pressure of 2.0×10^5 Pa at 57°C, what will the volume be in L?

____

(b) 0.32 g of oxygen was placed in a 500 mL container at 27°C. What is the pressure inside this container in Pa

____

2. Answer to the nearest whole number. Assume R = 8.3×10^3 Pa·L/(K·mol).

(a) When 1.2 g of a volatile substance was completely vaporized, its volume was 1.2 L at 77 °C and 1.0×10^5 Pa. Find the relative molecular mass.

(b) Find the relative molecular mass of a gas whose mass is 7.6 g for a volume of 4.15 L at 17 °C and 2.32×10^5 Pa.

(c) Find the relative molecular mass of a gas with a density of 2.4 g/L at 27 °C and 1.00×10^5 Pa.

Exercise

Choose the correct answer:

1. Why is the value of R the same for all ideal gases?

A. All gases have the same colour under the same conditions.

B. All gases have the same mass under the same conditions.

C. All gases behave the same way under the same conditions.

D. All gases have the same density under the same conditions.

2. A chemist measures the mass, pressure, volume, and temperature of an unknown gas and uses the Ideal Gas Law to calculate its molar mass. Which quantity must be measured in addition to P, V, and T?

A. Atomic radius

B. Gas colour

C. Mass of the gas sample

D. Density of water

Answer: ___________________________________

Answer on the Following Questions:

3. Assume R = 8.3×10^3 Pa·L/(K·mol).

(a) There is 0.10 mol of carbon dioxide with a volume of 2.9 L at 17 °C. What is its pressure in Pa?

Answer: ___________________________________

(b) When 2.0 mol of gas is subjected to 1.2×10^5 Pa at 27 °C, what will the volume be in L?

Answer: ___________________________________

(c) To expand 0.50 mol of gas to 33.2 L at 4.0×10^4 Pa, what should the temperature be in °C?

Answer: ___________________________________

(d) What is the amount of substance in mol of 16.6 L of gas at 127 °C and 3.0×10^5 Pa?

Answer: ___________________________________

Challenge Your Thinking

Designing an Investigation

A gas sample occupies 2.50 L at 27°C and 1.00×10^5 Pa.

The temperature is increased until the pressure doubles while the volume remains constant.

Without performing calculations:

- Predict what happens to the average kinetic energy of the gas particles. __________
- Explain how the particle collisions change. ___________________________________
- Justify your answers using the Ideal Gas Equation ___________________________

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

1-6 — Total Pressure and Partial Pressure

Engaging Introduction

Imagine a deep-sea diver breathing a special gas mixture, or an astronaut inside a spacecraft surrounded by carefully controlled gases. What will they feel in each case? Let's explore Dalton's law and how it explains why the pressure of each gas in a breathing tank increases with depth — affecting a diver's safety underwater.

❓ **Guiding question:** How can invisible gases share the same space but still behave as independent gases?

Exploration

(hook phenomena)

A scuba diver uses a mixture of oxygen and nitrogen instead of pure oxygen.Why?

- Oxygen is necessary for life.

- Nitrogen affects the total pressure.

- Changing the ratio of gases changes the pressure effects on the diver's body.

Point!

Total Pressure and Partial Pressure

(a) Total pressure: the pressure of a gas mixture (several gases that do not react with each other).

(b) Partial pressure: the pressure each component gas would exert if it alone occupied the same volume as the mixture.

(c) Dalton's law of partial pressures: Under a constant volume and temperature, the total pressure of a gas mixture is equal to the sum of the partial pressures of its component gases.

- If the total pressure of a gas mixture is P and the partial pressures of its component gases are P_1, P_2, P_3, ...,

then the following equation applies: $P = P_1 + P_2 + P_3 + ...$

⚠ Misconunderstanding Alert

✖ **Wrong idea:**

"After mixing gases, each gas loses its own pressure."

✓ **Correct:**

Each gas keeps its own partial pressure and contributes to total pressure.

✖ **Wrong idea:**

"The largest gas molecule creates the highest pressure."

✓ **Correct:**

Pressure depends mainly on the number of particles, not molecule size.

Structure of Science

Key Fact :

The behaviour of a gas mixture can be predicted by analysing the contribution of each component gas through its mole fraction and partial pressure.

Key Concepts:

- Total pressure
- Partial pressure
- Mole fraction
- Gas mixtures
- Average relative
- molecular mass
- Vapour pressure
- Atmospheric pressure.

Scientific law:

Dalton's law of partial pressures.

(d) Mole fraction: : The ratio of the amount of substance of each component gas to the total amount of substance of the gas mixture.

- If the amount of substance of gas A is nA [mol] and the amount of substance of gas B is nB [mol], the mole fraction of gas A in the mixture of gas A and gas B is $\frac{n_A}{n_A + n_B}$, and the mole fraction of gas B is $\frac{n_B}{n_A + n_B}$.
- The partial pressure of each component gas in a gas mixture can be expressed as:

Partial pressure = Total pressure × Mole fraction .

(e) **Relationship between partial pressure, volume, and amount of substance**

- At a constant temperature and volume, the ratio of the partial pressures of the component gases in a gas mixture is equal to the ratio of their amounts of substance.
- At a constant temperature and pressure, when each component gas is extracted individually, the ratio of the volumes of the component gases in the gas mixture is equal to the ratio of their amounts of substance.

(f) **Average relative molecular mass**

The apparent relative molecular mass of a mixture, assuming it consists of a single type of molecule, is the average relative molecular mass. It equals the sum of Relative molecular mass × Mole fraction of each component.

(g) **Partial pressure of a gas collected over water**

The gas collected over water is mixed with water vapour.

When the water levels inside and outside the collection container are equalized, the total pressure of the collected gas equals the atmospheric pressure.

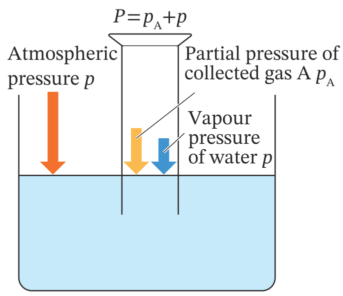


Figure 1.6.1. Measuring the atmospheric Pressure

As shown in the right figure 1.6.1, when gas A is collected over water, Atmospheric pressure = Total pressure of the collected gas = Partial pressure of gas A + Vapour pressure of water .

By rearranging this, we obtain: Partial pressure of gas A = Atmospheric pressure − Vapour pressure of water .

Equation box

Quantity	Relation
Dalton's law	$P = P_1 + P_2 + P_3 + \dots$
Mole fraction of A	x_A = nA / (nA + nB)
Partial pressure	Partial pressure = total pressure × mole fraction
Average relative molecular mass	Σ (relative molecular mass × mole fraction)
Collected over water	Partial pressure of gas = atmospheric − water-vapour pressure

★Testing Hypothesis (Scientific Inquiry)

Hypothesis:
„If the amount of gas increases inside a fixed-volume container, then its partial pressure will increase because more particles collide with the container walls."

Variables:
- Independent variable ___________________________________
- Dependent variable: ___________________________________
- Controlled variables: ___________________________________

Evidence Collection: ___________________________________

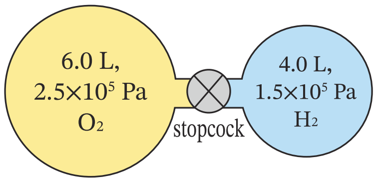
Investigation Design: ___________________________________

Warm Up

Answer the following questions.

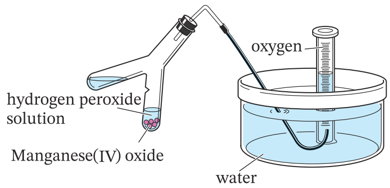
There is a 6.0 L container with oxygen at 2.5×10^5 Pa and a 4.0 L container with

hydrogen at 1.5×10^5 Pa at the same temperature.



Find the partial pressure of oxygen and the total pressure in the containers when the two containers are connected, the stopcock is opened, and kept at the same temperature for a sufficiently long time.

As shown in the figure on the right, oxygen was collected over water, and when the water levels inside and outside the measuring cylinder were equalized, its volume was 0.83 L at 1.03×10^5 Pa and 27°C. Answer the following questions with two significant figures. Assume the gas constant R = 8.3×10^3 Pa.L/(K.mol).



(i) Assuming the saturated vapour pressure of water at 27°C is 3.0×10^3 Pa, what is the partial pressure of the collected oxygen in Pa?

(ii) Find the amount of substance of the collected oxygen

Explanation

(a) To calculate the partial pressure of oxygen, consider the pressure when oxygen alone occupies the same volume as the gas mixture. When the two containers are connected, the volume of the gas mixture becomes 10.0 L.

Therefore, if the partial pressure of oxygen is P_A [Pa], from Boyle's law $P_1V_1 = P_2V_2$:

2.5×10^5 Pa × 6.0 L = P_A [Pa] × 10.0 L

$P_A = 1.5\times10^5$pa

Similarly, if the partial pressure of hydrogen is PB [Pa], from Boyle's law:

1.5×10^5 Pa × 4.0 L = P_B [Pa] × 10.0 L

$P_B = 6.0\times10^4$ Pa

Since the total pressure in the container is the sum of the partial pressures:

P_A [a]+P_B[Pa] = 1 .5× 10^5 Pa + 6.0 ×10^4 pa = 2.1× 10^5 pa

Therefore, the partial pressure of oxygen is 1.5×10^5 Pa, and the total pressure is 2.1×10^5 Pa.

(b) (i) If the partial pressure of oxygen is Px [Pa], from Partial pressure of oxygen = Atmospheric pressure – Vapour pressure of water:

$P_x = 1.03\times10^5 - 3.0\times10^3 =$ Answer: 1.0×10^5 Pa

(ii) From the ideal gas equation PV=nRT:

1.0×10^5 Pa × 0.83 L = n × 8.3×10^3 Pa.L/(K.mol) × (27+273) K

n = 3.33 ×10^{-2}

≈ 3.3×10^{-2} Answer: 3.3×10^{-2} mol

Try

1. As shown in the figure, nitrogen (relative molecular mass 28) at 2.0×10^5 Pa is placed in a 3.0 L container A, and hydrogen (relative molecular mass 2.0) at 1.0×10^5 Pa is placed in a 2.0 L container B. The stopcock is opened to mix the two gases.

The temperature was kept constant at all times. Answer the following questions about the mixed gas with two significant figures.

(a) What is the partial pressure of nitrogen in Pa?

Answer: ___________________________________

(b) What is the total pressure in Pa?

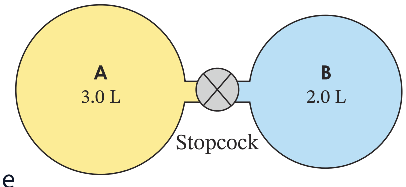
Answer: ___________________________________

(c) What is the mole fraction of each gas?

Answer: ___________________________________

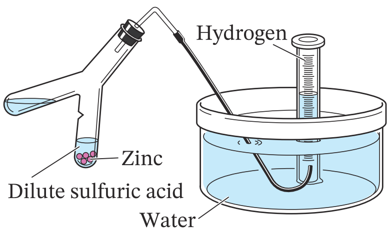
(d) What is the average relative molecular mass of the mixture?

Answer: ___________________________________



2. Hydrogen was collected over water, and when the water levels inside and outside the measuring cylinder were equalized, 498 mL of gas was obtained at 27°C and 1.04×10^5 Pa. What is the amount of substance in mol of the obtained hydrogen?

Answer with two significant figures. Assume the gas constant R = 8.3×10^3 Pa·L/(K·mol) and the vapour pressure of water at 27°C is 4.0×10^3 Pa.



Exercise

Choose the correct answer:

1. A sample of hydrogen is collected over water. The atmospheric pressure is 101 kPa and the water-vapour pressure is 3 kPa. What is the partial pressure of hydrogen?

A. 98 kPa B. 101 kPa

C. 104 kPa D. 3 kPa

Answer: ___________________________________

2. A student claims that increasing the amount of nitrogen in a mixture will increase the partial pressure of oxygen. Which statement best evaluates this claim?

A. Correct, because all partial pressures increase together.

B. Correct, because total pressure stays constant but molar mass increases.

C. Incorrect, because oxygen's partial pressure depends on its own mole fraction.

D. Incorrect, because nitrogen does not contribute to pressure.

Answer: ___________________________________

3. Answer the following questions with two significant figures.

(a) Maintaining the same temperature, gas A at 6.0×10^4 Pa, 250 mL and gas B at 4.0×10^4 Pa, 150 mL were placed in a 500 mL container. Answer the following questions.

(i) Find the partial pressure of gas A.

____

(ii) Find the partial pressure of gas B.

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

____

(iii) Find the total pressure of the gas mixture.

____

(b) At a constant temperature, when the pressure of a gas mixture of 4.8 g of methane CH_4 and 2.8 g of carbon monoxide CO is 1.0×10^5 Pa, what is the partial pressure of methane in Pa? (H = 1.0, C = 12, O = 16)

____

4. Oxygen was generated by adding manganese(IV) oxide to hydrogen peroxide solution, and it was collected with a

measuring cylinder over water under an atmospheric pressure of 1.0×10^5 Pa at 27°C. When the water levels inside and outside the measuring cylinder were equalized, the volume of the collected gas was 700 mL. Find the amount of substance of the oxygen generated at this time with two significant figures. Assume the saturated vapour pressure of water at 27°C is 3.6×10^3 Pa and the gas constant is 8.3×10^3 Pa·L/(K·mol)

🤔 **Think Critically** — A scientist analyzes air from two different locations:

- Location A: High mountain area
- Location B: Sea level

Both samples contain oxygen and nitrogen.

1. if the percentage of oxygen is almost the same in both locations, but the total pressure is different. Why does the oxygen partial pressure change?

____

2. If the amount of oxygen stays constant but nitrogen is added, what happens to:

- Total pressure?_______________________________
- Oxygen partial pressure?_______________________________

1-7 — Ideal Gas and Real Gas

✨ Engaging Introduction

A balloon left outside on a cold morning becomes smaller, but when placed in sunlight it expands again. Let›s explore why the gas inside behaves differently when the temperature changes.

> ❓ **Guiding question:** Why do real gases deviate from ideal behaviour, and how can scientists explain these deviations?

Exploration

Consider the following situations:

• Gas stored in a scuba tank at extremely high pressure.

• Liquid nitrogen slowly warming to room temperature.

• Carbon dioxide compressed in a fire extinguisher.

• Natural gas transported through long pipelines.

Discuss:

• Would you expect gases to behave ideally in these situations?

____

• Which conditions might cause deviations?

____

• What evidence would help you test your prediction?

____

structure of science

Key Fact :

An ideal gas is a simplified model in which gas particles have no volume and no intermolecular forces.

Key Concepts:

Ideal gas · Real gas · Intermolecular forces · Molecular volume · Compression factor (Z)

Scientific principles:

intermolecular attractions reduce the observed pressure; finite molecular volume becomes important at high pressures.

Point!

The ideal gas equation of state, PV = nRT, does not perfectly describe all gases. Real gases deviate from ideal behaviour because of the finite size of their molecules and the intermolecular forces between them.

A. Definitions of ideal gases and real gases

- Ideal Gas: a hypothetical model assumed to simplify calculations.

(1) The volume of the gas molecules themselves is zero.

(2) No intermolecular forces act between the molecules.

- Real Gas: gases that actually exist, such as hydrogen or nitrogen.

(1) The molecules have a small, finite volume.

(2) Attractive (intermolecular) forces act between the molecules.

B. Why do deviations occur?

The deviation of a real gas from ideal behaviour is quantified by the Compression Factor (Z):

$$Z = \frac{PV}{nRT}.$$

For an ideal gas, Z is always equal to 1. However, for a real gas, the value of Z varies depending on the conditions of **pressure** and **temperature**. This variation is related to the following two physical factors.

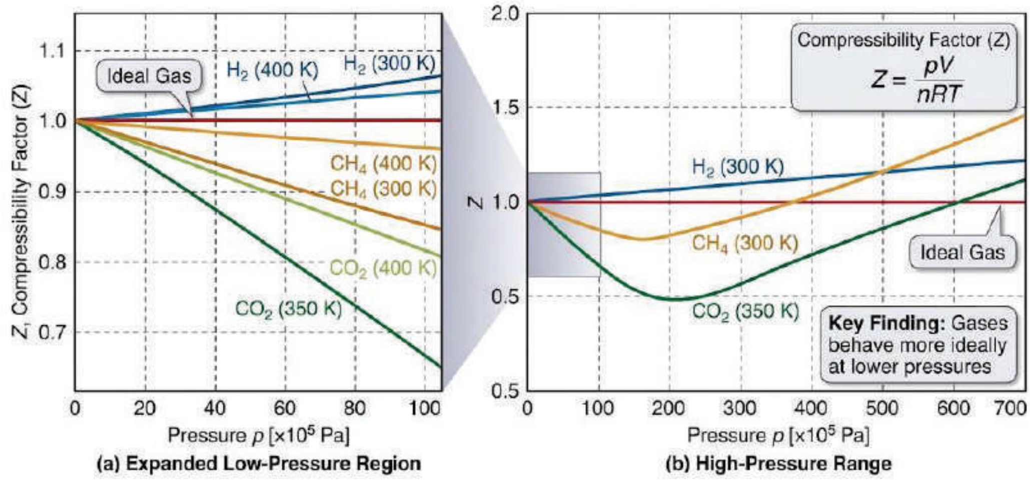
(i) When attractive forces act between molecules, the force with which the molecules collide with the container walls is weakened. As a result, the **observed pressure, P, is lower than that predicted for an ideal state**.

This effect is pronounced under conditions of **moderate pressure** (where molecules are closer together) and **low temperature**, causing Z to be less than 1.

(ii) Effect of Molecular Volume (Reason for Z>1)

At very high pressures, the **volume of the molecules themselves** is no longer negligible compared to the total volume, V, of the gas. Since the **free volume available for molecular motion effectively decreases**, the gas becomes harder to compress further.

Consequently, the observed volume, V, is greater than the calculated ideal value, and Z significantly exceeds 1 under these conditions.



1-7-1 Compressibility Factor (Z) vs. Pressure for Real Gases

C. Conditions for approaching ideal-gas behaviour

(i) **Effect of pressure:**

When a real gas is compressed, increasing the pressure, the number of molecules per unit volume increases, and thus the proportion of volume occupied by the molecules themselves becomes significant. Due to the influence of this molecular volume, the actual volume of the gas becomes greater than that of an ideal gas.

When the pressure is lowered, the number of molecules per unit volume decreases. This minimises the effect of molecular volume, causing the real gas to approach ideal behaviour.

(ii) **Effect of temperature:**

As the temperature of a real gas is lowered, the thermal motion of the molecules slows down, and the effect of intermolecular forces becomes greater. This results in the actual pressure of the gas being lower than that of an ideal gas.

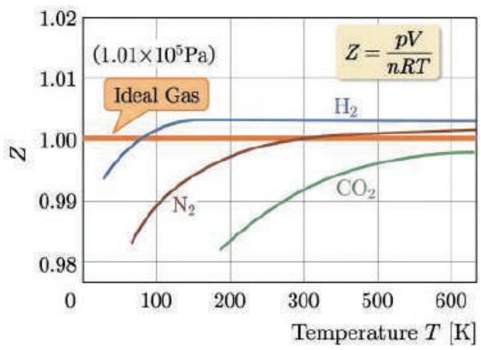


Figure 1.7.2 the effect of temperature on Ideal gas

Graph Literacy

Reading the Compression Factor Graph

Before reading the explanation:

- Where is Z equal to 1?

……………………………

- Where is Z less than 1?

……………………………

- Where is Z greater than 1?

……………………………

- Which region suggests stronger intermolecular attractions?

……………………………

- Which region suggests the effect of molecular volume?

……………………………

Graph Detective

Use evidence from the graph to support your answers.

……………………………

……………………………

Conversely, raising the temperature increases the kinetic energy (thermal motion) of the molecules, which minimises the effect of intermolecular forces, causing the real gas to approach ideal behaviour.

Therefore, a real gas can be treated as an ideal gas under conditions of high temperature and low pressure.

Think as an Engineer

Engineers who design:

- Industrial gas tanks
- Deep-sea diving systems
- Spacecraft fuel systems

cannot assume perfect gas behavior under all conditions. Ignoring real-gas effects could lead to inaccurate predictions and unsafe designs.
Why is understanding model limitations important in engineering?

____

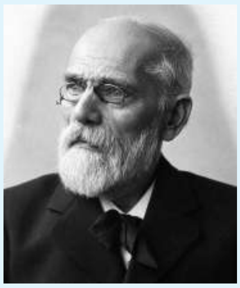
____

D. [Advanced] A more accurate model: the Van der Waals equation of state

The ***Van der Waals equation of state*** *was developed to account for real-gas deviations. It adds a correction term for intermolecular forces, a, and one for molecular volume, b:*

$$\left(P+\frac{n^2}{V^2}a\right)(V-nb)=nRT$$

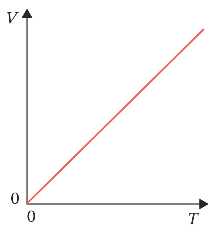
Appreciating Scientists



While many scientists used the Ideal Gas Model successfully, Johannes Diderik van der Waals recognized that real gases possess molecular volume and intermolecular attractions.

His work led to the development of a more realistic equation of state and transformed our understanding of matter.

E. Ideal Gas And The Ideal Gas Equation

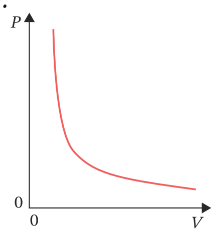


For an ideal gas, since the ideal gas equation PV = nRT holds, the relationship between pressure, volume, amount of substance, and temperature can be determined.

Example 1 *— Relationship between volume and temperature when the pressure of a given amount of gas is constant In PV = nRT, P and n can be regarded as constants.*

$$\text{Solving for } V,\ V=\underbrace{\frac{nR}{P}}_{\text{Constant}}T$$

Therefore, the volume V is directly proportional to the temperature T.



Example 2 *— Relationship between volume and pressure when the temperature of a given amount of gas is constant*

In PV = nRT, n and T can be regarded as constants.

$$\text{Solving for } V,\ V=\frac{\overbrace{nRT}^{\text{Constant}}}{P}$$

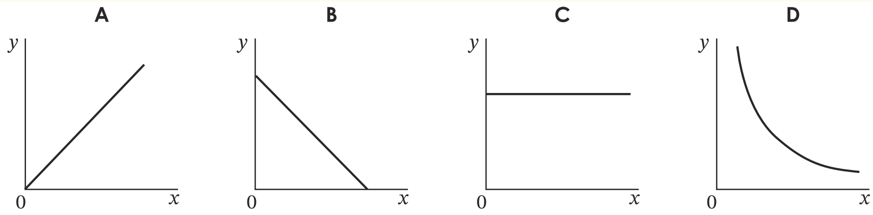
Therefore, the volume V is inversely proportional to the pressure P.

Key relations

Quantity	Meaning
Z = PV / nRT	compression factor; Z = 1 for an ideal gas
Z < 1	intermolecular forces dominate (moderate P, low T)
Z > 1	molecular volume dominates (very high P)
Van der Waals	$(P + n^2a / V^2)(V - nb) = nRT$
Near-ideal conditions	high temperature and low pressure

Warm Up

Regarding an ideal gas, choose the graph (A to D) that represents each relationship (a) to (c).



(a) The relationship between volume x and pressure y when the amount of substance and temperature of the gas are constant.

____

(b) The relationship between the amount of substance x and pressure y when the temperature and volume are constant.

____

(c) The relationship between pressure x and the product of pressure and volume y when the amount of substance and temperature of the gas are constant

____

Explanation

By rearranging the ideal gas equation, PV = nRT.

(a) Solving the ideal gas equation for pressure to obtain an equation in terms of y, $P = \frac{nRT}{V}$

*When the amount of substance and temperature of the gas are constant, nRT is a constant, so the pressure P is inversely proportional to the volume **V**.*

Therefore, the graph showing the relationship between x and y is D.

(b) Solving the ideal gas equation for pressure to obtain an equation in terms of y, $P = \frac{nRT}{V}$

When the temperature and volume are constant, $\frac{RT}{V}$ *is a constant, so the amount of substance n is directly proportional to the pressure P.*

*Therefore, the graph showing the relationship between x and y is **A**.*

(c) When the amount of substance and temperature of the gas are constant, the right side of PV = nRT remains constant. ● The left side PV is y.

*Regardless of the value of pressure P, the product of pressure and volume PV is always constant, so the graph showing the relationship between x and y is **C**.*

Try

Choose the correct answer:

1. *Which statement correctly describes an ideal gas?*

A. Gas molecules have finite volume and attractive forces.

B. Gas molecules have zero volume and no intermolecular forces.

C. Gas molecules have large volume but no attractive forces.

D. Gas molecules occupy half of the container volume.

Answer: ______________________________________

2. *Under which conditions will a real gas behave most like an ideal gas?*

A. Low temperature and high pressure.

B. High temperature and low pressure.

C. Low temperature and low pressure.

D. High temperature and high pressure.

Answer: ______________________________________

3. *Regarding ideal gases and real gases, select all the correct statements from (i) to (iv) below.*

(i) "A gas where the molecules themselves have no volume, and the molecules move freely without intermolecular forces

acting between them" is a real gas.

(ii) For a real gas, the ideal gas equation holds strictly at any temperature, volume, and pressure.

(iii) As the temperature rises, the deviation of a real gas from an ideal gas becomes smaller.

(iv) Comparing hydrogen at 300 K and 1.0×10^5 Pa with hydrogen at 350 K and 4.0×10^4 Pa, the hydrogen at 300 K and 1.0×10^5 Pa is closer to an ideal gas.

Exercise

Regarding an ideal gas, choose the graph (A to E) that represents each relationship (a) to (e). You may choose the same letter multiple times.

Answer: ______________________________________

(a) The relationship between the amount of substance x and the volume y of a gas when the temperature and pressure are constant.

Answer: ______________________________________

(b) The relationship between the absolute temperature x and the volume y of a gas when the amount of substance and pressure are constant.

Answer: ______________________________________

(c) The relationship between the absolute temperature x and the pressure y of a gas when the amount of substance and volume are constant.

Answer: ______________________________________

(d) The relationship between the pressure x and the volume y of a gas when the amount of substance and temperature are constant.

Answer: ______________________________________

(e) The relationship between the pressure x and the product of pressure and volume y of a gas when the amount of substance and temperature are constant.

Answer: ______________________________________

Nature of Science

Scientific knowledge evolves over time.

When observations could not be fully explained by the Ideal Gas Model, scientists developed improved models such as the Van der Waals Equation.

Reflect and Grow

Complete:

✓ One idea that changed my thinking

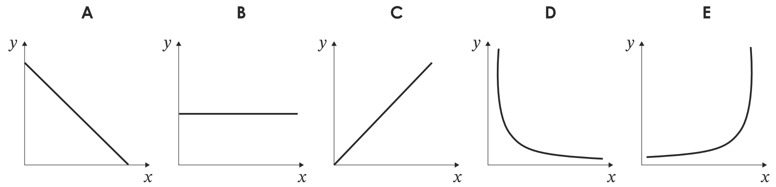
....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....



Answer: ___________________________________

Challenge Your Thinking

A real gas is compressed to an extremely high pressure while its temperature remains high.

Predict:

• Will the gas behave more ideally or less ideally?

____

• Which factor is likely to dominate: intermolecular forces or molecular volume?

____

• Will Z most likely be greater than 1, less than 1, or close to 1?

____

Justify your answer using particle-level reasoning and evidence from the lesson.

____

1-8 — Properties of Solutions

Engaging Introduction

A student vigorously shakes a bottle of water and cooking oil. For a short time the liquids appear mixed, but a few minutes later two separate layers form again. Many people assume that stirring longer should keep them mixed.

Let›s explore how substances dissolve, and what factors decide whether they stay dissolved or separate from a solution.

> ❓ **Guiding question:** Why do some substances dissolve easily while others do not, and how can scientists predict and control solubility?

Exploration

Predict Before You Learn

Without conducting an experiment, predict what would happen if you add the following substances to water:

• Table salt

• Sugar

• Vegetable oil

• Candle wax

For each substance:

• Will it dissolve? ___________________________________

• Why do you think so? ___________________________________

• What evidence supports your prediction?

Point!

1. Dissolution

(a) **Dissolution:** The phenomenon where a substance dissolves uniformly in a solvent.

(b) **Many ionic crystals and dissolution:**

Ionic crystals readily dissolve in water and are largely insoluble in non-polar solvents.

(c) **Polarity and dissolution:**

Substances with similar polarity tend to dissolve in one another.

• Polar molecules readily dissolve in polar solvents. (Because the positive and negative charges attract each other).

• Non-polar molecules readily dissolve in non-polar solvents.

• Substances composed of polar molecules and non-polar molecules do not readily dissolve in each other.

(d) **Solubility:** The maximum amount of a solute that dissolves in a given amount of solvent. In the case of solids, it is often expressed as the maximum mass of solute that dissolves in 100 g of solvent.

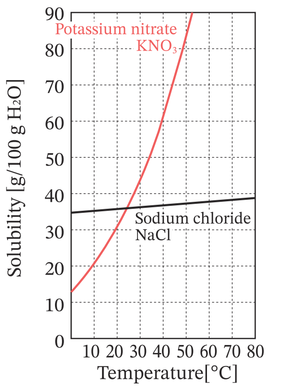


Fig 1.8.1 Solubility Curve

structure of science

Key Fact :

The solubility of a substance depends on the nature of the solute and solvent — substances with similar polarity tend to dissolve in one another.

Key Concepts:

- Dissolution
- Solubility
- Solvent and solute
- Polarity
- Ionic crystals
- Concentration
- Molarity
- Molality
- Recrystallization

Scientific law:

Henry›s law (gas solubility ∝ gas

(e) **Solubility curve:** A curve showing how solubility varies with temperature (Figure 1.8.1).

(f) **Recrystallization:** Recovering a dissolved substance as solid crystals by utilizing the difference in solubility due to temperature.

2. Concentration

(a) **Percentage by mass:** A concentration expressing the ratio of the mass of the solute to the mass of the solution.

Percentage by mass (%) = (mass of solute (g) / mass of solution (g)) × 100.

(b) **Molarity (molar concentration):** A concentration expressing the amount of solute dissolved in 1 L of solution in terms of amount of substance (mol).

Molarity [mol/L] = Amount of Substance of Solute (mol) / Volume of Solution (L).

(c) **Molality:** A concentration expressing the amount of solute dissolved in 1 kg of solvent in terms of amount of substance (mol).

Molality (mol/kg) = Amount of Substance of Solute (mol) / Mass of Solvent (kg).

Example — Molality of an aqueous solution in which 2.0 g of sodium hydroxide is dissolved in 23.0 g of water (NaOH = 40)

The amount of substance of sodium hydroxide is $\frac{2.0\ \text{g}}{40\ \text{g/mol}} = 0.050$ mol

Since the mass of water is 23.0 g = 0.023 kg,

the molality is: $\frac{0.050\ \text{mol}}{0.023\ \text{kg}} = 2.17 \approx 2.2$ mol/kg

3. Solubility of gases

Henry's law: At a constant temperature, the amount of substance and mass of a gas that dissolve in a given amount of solvent are directly proportional to the pressure of the gas.

• When considering the solubility of a gas, Henry's law is used.

Equation box

Concentration	Definition
Percentage by mass [%]	(mass of solute / mass of solution) × 100
Molarity [mol/L]	amount of solute [mol] / volume of solution [L]
Molality [mol/kg]	amount of solute [mol] / mass of solvent [kg]
Henry›s law	dissolved amount (mol or g) ∝ gas pressure

Chemistry in Everyday Life

Why Do Soft Drinks Fizz?

Soft drinks contain carbon dioxide dissolved under high pressure.

When the bottle is opened:

• Pressure decreases.

• Gas solubility decreases.

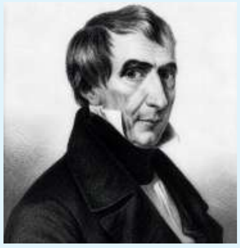
• Carbon dioxide escapes as bubbles.

This is a direct application of Henry's Law.

Appreciating Scientists

William Henry discovered an important relationship between gas pressure and gas solubility.

His work continues to influence environmental science, medicine, and industry today.



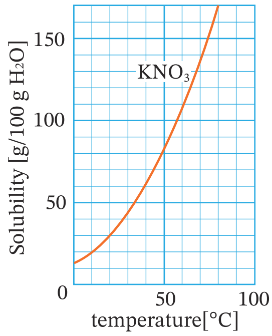
Warm Up

Answer the following questions.

(a) Among the following, choose all that dissolve in water but are largely insoluble in hexane:

A Naphthalene · B Glucose · C Potassium chloride.

(b) The figure on the right is the solubility curve of potassium nitrate KNO3.



Answer the following questions.

(i) When an aqueous solution saturated with potassium nitrate in 50 g of water at 50°C is cooled to 20°C, how many grams of crystals will crystallize out?

(ii) When 100 g of a saturated aqueous solution of potassium nitrate at 60°C is cooled to 20°C, how many grams of crystals will crystallize out?

(c) At 20°C and 1.0×10^5 Pa, 125 mL of oxygen dissolves in 1.0 L of water, converted to the volume at 0°C and 1.0×10^5 Pa.

What is the mass of oxygen in grams that dissolves in 1.5 L of water at 20°C and

2.0×10^5 Pa? Assume the volume of 1 mol of gas at 0°C and 1.0×10^5 Pa is 22.4 L. (O = 16)

Explanation

(a) Water is a polar solvent, and hexane is a non-polar solvent.

A. Since naphthalene is a non-polar molecule, it is largely insoluble in water and readily dissolves in hexane.

B. Since glucose is a polar molecule, it readily dissolves in water and is largely insoluble in hexane.

C. Since potassium chloride is an ionic crystal, it readily dissolves in water and is largely insoluble in hexane.

From the above, **B** and **C**.

(b) (i) From the figure, 84 g of potassium nitrate dissolves in 100 g of water at 50°C, so 42 g dissolves in 50 g of water.

Also, 30 g dissolves in 100 g of water at 20°C, so 15 g dissolves in 50 g of water. Therefore, the mass of crystals that crystallize out is 42−15 = 27 g..

(ii) From the figure, 108 g of potassium nitrate dissolves in 100 g of water at 60°C to form a saturated solution. The mass of this solution is 108 g+100 g = 208 g.

When this solution is cooled to 20°C, the mass of the precipitate is 108 g−30 g = 78 g.

Let x be the mass in grams of the precipitate formed from 100 g of the saturated solution.

Be a Global Citizen

Understanding solubility helps address global challenges such as water purification, ocean conservation, sustainable agriculture, and pollution control. Chemistry provides tools for making informed decisions about resources and environmental stewardship.

We can set up the following proportion: 100 g : x [g] = 208 g : 78 g

x = 37.5 ≈ 38 g Answer: 38 g

(c) Since only the pressure and volume of the dissolved gas are known, first calculate the moles of gas to apply Henry's Law.

At 0°C and 1.0×10^5 Pa, 1 mol of gas occupies 22.4 L.

Let x be the moles of 125 mL (0.125 L) of oxygen:

1 mol : 22.4 L = x [mol] : 0.125 L

$$x = 5.580\times10^{-3} \approx 5.58\times10^{-3} \text{ mol}$$

According to Henry's law, the solubility of a gas at a constant temperature is directly proportional to its partial

pressure. Let y be the moles of oxygen dissolved in 1.0 L of water at 20°C and 2.0×10^5 Pa:

1.0×10^5 Pa : 5.58×10^{-3} mol = 2.0×10^5 Pa : y [mol]

$$y = 1.11\times10^{-2} \text{ mol}$$

Let z be the moles of oxygen dissolved in 1.5 L of water:

1.0 L : 1.11×10^{-2} mol = 1.5 L : z [mol]

$$z = 1.66\times10^{-2} \text{ mol}$$

Thus, the required mass of oxygen is: 1.66×10^{-2} mol × 32 g/mol = 0.531 g ≈ Answer: 0.54 g

Try

Choose the correct answer:

1. Glucose dissolves readily in water but is largely insoluble in hexane. Which explanation best accounts for this?

A. Glucose and water are non-polar.

B. Water and glucose have similar polarity.

C. Hexane and water form an ionic solvent.

D. Water contains ions.

Answer: ____________________________________

2. A student cools a saturated solution of potassium nitrate and observes crystals forming. Which process is occurring?

A. Neutralization B. Sublimation

C. Recrystallization D. Ionization

Answer: ____________________________________

Answer on the Following Questions

3. The figure shows the relationship between the solubility of potassium nitrate KNO_3 and temperature. Answer the following questions.

(a) How many grams of potassium nitrate can dissolve in 500 g of water at 50°C? Answer with two significant figures.

____

(b) When a solution of 80 g of potassium nitrate dissolved in 100 g of water at 60°C is

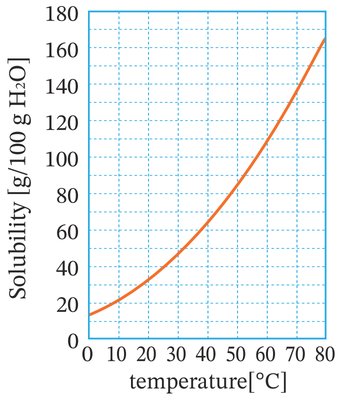
cooled to 20°C, how many grams of potassium nitrate crystals will be obtained?
Answer with two significant figures.

____

(c) When an aqueous solution of 10 g of potassium nitrate dissolved in 50 g of water

at 30°C is cooled, at what temperature do the crystals begin to crystallize out?

____



(d) When 100 g of a saturated solution at 50°C is cooled to 20°C, how many grams of potassium nitrate crystals will be obtained? Answer with two significant figures

____

Experimental Investigation

Which substance dissolves most effectively in water and which dissolves most effectively in a nonpolar solvent?

Materials

- Water
- Vegetable oil
- Salt
- Sugar
- Candle wax shavings

____

Exercise

Answer on the Following Questions

1. Among (i) to (iii) below, choose all the substances that dissolve in hexane but do not dissolve in water.

(i) Potassium nitrate (ii) Hydrogen chloride (iii) Naphthalene

Answer: ____________________________________

2. The figure shows the solubility curves of KNO_3 and $NaNO_3$.

(a) 50 g of water was saturated with potassium nitrate at 70°C. How many grams of potassium nitrate are dissolved in the aqueous solution? Answer with two significant figures.

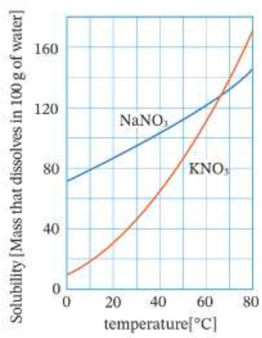
____

____

(b) When the solution from (a) is cooled to 20°C, how many grams of potassium nitrate will crystallize out? Answer with two significant figures.

____

____



Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

(c) 50 g of sodium nitrate and 20 g of potassium nitrate were dissolved in 50 g of water at 80°C. When this solution is cooled, which substance crystallizes out first? Answer with its chemical formula.

____

____

Challenge Your Thinking

A diver descends to a great depth where pressure is much higher than at the surface.

Without calculations:

• Predict what happens to the amount of nitrogen dissolved in the diver›s blood.

____

• Predict what might happen if the diver ascends too quickly.

____

• Explain your reasoning using Henry›s Law.

____

1-9 — Boiling Point Elevation and Freezing Point Depression

✨ Engaging Introduction

During winter, lakes in cold regions may freeze while nearby seawater often stays liquid at the same temperature. Both contain water, yet they behave differently. Scientists discovered that dissolved substances can change the temperatures at which liquids freeze and boil — with important consequences for oceans, climate, food preservation, and many industrial processes.

Let's explore how dissolved particles can dramatically alter the physical behaviour of liquids.

❓ Guiding question: How can dissolved particles change the temperature at which a liquid boils or freezes?

Exploration

Imagine adding increasing amounts of salt to water.

Predict what will happen to:

• The boiling point.

• The freezing point.

• The vapor pressure.

Explain your reasoning before looking at the graphs.

____

Point!

1. Vapour-pressure lowering and boiling point elevation

(a) **Vapour-pressure lowering:**

The phenomenon where the vapour pressure of a solution in which a non-volatile substance is dissolved in a solvent becomes lower than the vapour pressure of the pure solvent. (Figure on the right 1.9.1)

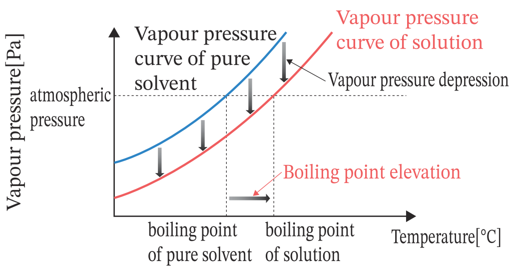


Figure 1.9.1 vapour pressure lowering curve

Reason: Because solute particles at the liquid surface inhibit the evaporation of solvent molecules.

(b) Boiling point elevation: The phenomenon where the boiling point of a solution becomes higher than the boiling point of the pure solvent.

(c) **Molal boiling point elevation constant:**

The boiling point elevation when the molality of the solution is 1 mol/kg.
It is represented by the symbol K_b.

structure of science

Key Fact :

Boiling point elevation and freezing point depression depend on the number of dissolved particles in a solution, not on the type of solute.

Key Concepts:

- Boiling point - elevation
- Freezing point depression
- Supercooling ·
- Non-volatile solute ·
- Cooling curve.

Scientific law:

Boiling Point Elevation Δt=Kbm

Freezing Point Depression Δt=Kfm

Think Like a Chemist

A chemist wants to recover pure crystals from a dissolved substance.

Instead of evaporating all the solvent, the chemist cools the solution.

Reflection

Why can cooling be an effective purification method?

____

Connect your answer to solubility and crystallization.

____

2. Freezing point depression

(a) **Cooling curve:** A graph showing the relationship between the temperature of a substance and the cooling time. (Figure on the right 1.9.2)

(b) **Freezing point:** The temperature at which the solvent in a solution begins to freeze.

(c) **Freezing point depression:**

The phenomenon where the freezing point of a solution in which a non-volatile substance is dissolved in a solvent becomes lower than the freezing point of the pure solvent.

(d) **Molal freezing point depression constant:** The freezing point depression when the molality of the solution is 1 mol/kg. It is represented by the symbol Kf.

(e) **Supercooling:** The phenomenon where the temperature of a liquid falls below its freezing point while remaining in a liquid state as it is cooled.

• When freezing begins in a supercooled state, the temperature rises.

• The cooling curve of a solution shows that the temperature continues to fall even as freezing progresses.

Reason: Because only the solvent freezes, the concentration of the solution increases, and the freezing point depression becomes larger.

• When supercooling occurs, the freezing point is taken as the temperature at the point where the horizontal plateau line (extended to the left) intersects the original cooling curve. (see Figure above 1.9.2)

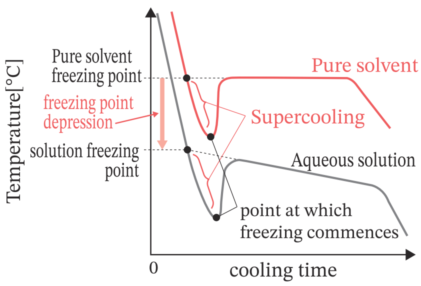


Figure 1.9.2. Freezing point depression

Graph Literacy

Examine the Freezing point depression graph carefully.

Identify:

• Where super-cooling occurs?

____

• Where freezing begins?

____

• The actual freezing point.

____

• Evidence that freezing is occurring.

____

Explain

Why does the temperature rise slightly after super-cooling?

____

3. Magnitude of boiling point elevation and freezing point depression

(a) Magnitude of boiling point elevation and freezing point depression

(i) It is independent of the type of solute.

(ii) It is directly proportional to the molality of the solution.

Note: When the solute is an electrolyte, the molality of the solution is considered to be directly proportional to the number of particles in the solution.

Example: Sodium chloride NaCl ionizes into Na^+ and Cl^- in aqueous solution, and the number of particles in the solution doubles.

⇒ The molality of the solution is considered to be doubled.

(b) Magnitude of molal boiling point elevation and molal freezing point depression constants

(i) They take values determined by the solvent, regardless of the type of solute.

(ii) It serves the proportionality constant K relating the magnitude of the boiling point elevation or freezing point depression and the molality.

- When the molality of the solution is m [mol/kg] and the magnitude of the boiling point elevation or freezing point depression is Δt,

$\Delta t = Km$ (K is the molal boiling point elevation constant Kb or the molal freezing point depression constant Kf), (m: molality of dissolved molecules and ions)

📦 Equation box

Quantity	Relation
Boiling point elevation	$\Delta t = K_b \cdot m$
Freezing point depression	$\Delta t = K_f \cdot m$
Electrolyte solute	multiply molality by the number of ions per formula unit

🚗 Chemistry around us

Antifreeze in Car Radiators

In cold climates, water inside a car's radiator could freeze when temperatures drop below 0°C. If this happens, the expanding ice may damage the cooling system and engine components.

To prevent this problem, a substance such as ethylene glycol is added to the radiator water. The dissolved particles lower the freezing point of the solution, allowing it to remain liquid at temperatures where pure water would freeze.

Warm Up

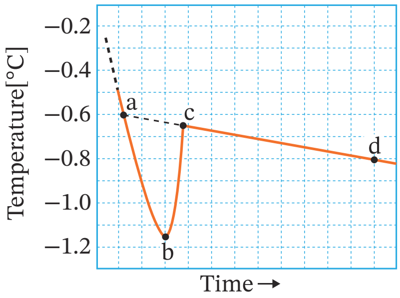
Answer the following questions.

(a) The figure shows the temperature change when a solution of non-electrolyte A dissolved in 100 g of water is cooled. Answer the following questions.

(i) What is the cooling state between a and b in the figure called?

(ii) At which point, a, b, or c, do crystals begin to form?

(iii) Which point's temperature from a to d is the freezing point of the solution?



(b) Answer the following questions. Assume the freezing point of water is 0°C, and the molal freezing point depression constant of water is 1.85 K·kg/mol.

(i) Find the freezing point of an aqueous solution in which 11.7 g of sodium chloride NaCl (formula mass 58.5) is dissolved in 500 g of water. Assume sodium chloride is completely ionized.

(ii) The difference in freezing point between an aqueous solution of 7.5 g of a certain non-volatile non-electrolyte dissolved in 250 g of water and pure water was 0.30°C. At this time, find the relative molecular mass of the non-electrolyte as an integer.

Explanation

(a) (i) Supercooling

(ii) When freezing begins in a supercooled state, the temperature rises, so the lowest temperature point is the point where freezing begins and crystals begin to form. b

(iii) a

The freezing point is where line segment cd intersects the temperature axis when extended to the left.

(b) (i) When finding the freezing point of an aqueous solution, first find the magnitude of the freezing point depression of the aqueous solution.

Since the freezing point of water is 0°C, the freezing point of an aqueous solution can be determined once the magnitude of the freezing point depression is known.

If the required magnitude of the freezing point depression of the aqueous solution is Δtf, the molal freezing point depression constant is Kf, and the molality of the solution is m [mol/kg], the following equation applies: Δtf = Kfm

Therefore, first find the molality m of the solute particles.

The amount of substance of the dissolved sodium chloride is $\frac{11.7\ \text{g}}{58.5\ \text{g/mol}}$ = 0.200 mol, and the mass of water is 500 g = 0.500 kg, so

m = $\frac{0.200\ \text{mol}}{0.500\ \text{kg}}$ = 0.400 mol/kg

Here, the sodium chloride NaCl aqueous solution ionizes to produce Na+ and Cl−, and the number of particles doubles, so the molality is also considered to be doubled. Therefore,

0.400 mol/kg × 2 = 0.800 mol/kg

Thus, the magnitude of the freezing point depression is:

Δtf = 1.85 K.kg/mol × 0.800 mol/kg = 1.48 K

Since the freezing point of water is 0°C, the required freezing point of the aqueous solution is:

$0 - 1.48 = -1.48°C$

(ii) Since the molal freezing point depression constant is given, the relative molecular mass is found using the molality m [mol/kg].

Let the molar mass of the non-electrolyte be M [g/mol], the molality is: $\dfrac{\frac{7.5}{M}\,[\text{mol}]}{0.250\ \text{kg}}$

Also, the difference from the freezing point of pure water is the magnitude of the freezing point depression of the aqueous solution.

If the magnitude of the freezing point depression is Δtf and the molal freezing point depression constant is Kf, then $\Delta tf = Kfm$ holds, so

$$0.30\ \text{K} = 1.85\ \text{K·kg/mol} \times \dfrac{\frac{7.5}{M}\,[\text{mol}]}{0.250\ \text{kg}}$$

Solving this gives M = 185 g/mol. Therefore, the relative molecular mass is 185.

Think Critically

Evidence or Assumption?

Two containers are placed in a freezer:

- Container A contains pure water.
- Container B contains salt water.

Container A freezes first.

A student concludes:

«The freezer cooled Container A more effectively.»

Is this conclusion supported by evidence? ....

What alternative explanation could account for the observation? ....

Try

Choose the correct answer:

1. A scientist doubles the molality of a solution while keeping the solvent unchanged. What happens to the boiling point elevation?

A. It remains unchanged.

B. It doubles.

C. It is halved.

D. It becomes zero.

Answer: ___________________________________

2. A student cools a saturated solution of potassium nitrate and observes crystals forming. Which process is occurring?

A. Neutralization

B. Sublimation

C. Recrystallization

D. Ionization

Answer: ___________________________________

Be a Global Citizen

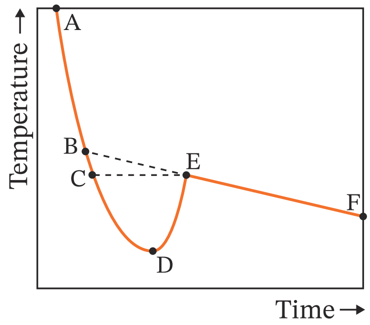
Climate, Ice, and Transportation

Countries around the world use salts to reduce accidents caused by icy roads.

However, excessive use of road salts can affect soil quality, freshwater ecosystems, and drinking-water supplies.

Solve the Following Questions

3. The figure is a cooling curve showing the relationship between cooling time and temperature when a dilute aqueous solution of sucrose $C_{12}H_{22}O_{11}$ is cooled. Answer the following questions.



(a) What is the phenomenon where the temperature drops while remaining a liquid

even after passing the freezing point, as shown in the figure? Write its name.

____

(b) Which point's temperature from A to F in the figure is the freezing point?

____

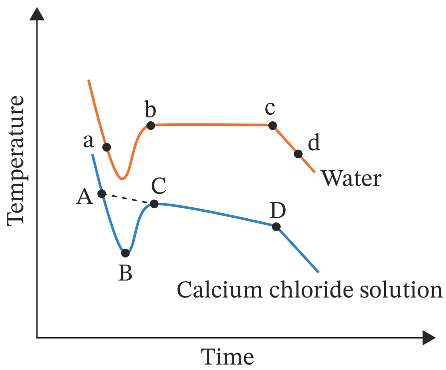
(c) Explain why the straight line EF in the figure slopes downward to the right

____

Exercise

Answer on the Following Questions

The figure on the right is a cooling curve showing the relationship between cooling time and temperature when water and a calcium chloride aqueous solution are slowly cooled. Answer the following questions.



(a) What is the phenomenon occurring to the water at a?

____

(b) Which point (A to D) in the figure represents the freezing point of the calcium chloride aqueous solution?

____

(c) Where does the freezing of the calcium chloride aqueous solution begin? Choose from A to D in the figure.

____

(d) In the range from b to c, the temperature is constant despite being cooled. Briefly explain the reason.

____

(e) In the range from C to D, the temperature gradually drops. Briefly explain the reason

____

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

Grand Challenge

Designing an Eco-Friendly De-Icer

Your city wants to reduce winter road accidents while minimizing environmental damage.

You have access to:

- Sodium chloride
- Calcium chloride
- Sugar
- Sand

Using your understanding of freezing point depression:

- Which material would you recommend? ....
- What evidence supports your choice? ....
- What environmental trade-offs should be considered? ....

1-10 — Osmotic Pressure

Engaging Introduction

A cucumber is placed in a concentrated salt solution and left for several hours. Later it becomes noticeably smaller and softer. Nothing seems to have been taken from the cucumber, yet its size has changed. **Where did the water go?**

In this lesson we will examine how careful observations of water movement in living organisms led scientists to develop models and explanations for osmosis and osmotic pressure.

> ❓ **Guiding question:** How and why does a solvent move across a semipermeable membrane, and how can scientists measure the pressure this produces?

Exploration

Investigate Water Movement

Imagine a U-tube separated by a semipermeable membrane.

One side contains pure water.

The other side contains a glucose solution.

Without studying the lesson:

• Predict the direction of water movement.

____

• Predict what will happen to the liquid levels.

____

• Explain your reasoning.

____

Point!

Osmosis and osmotic pressure

(i) **Semipermeable membrane:**

A membrane with tiny pores that lets small solvent molecules (e.g. water) pass but blocks large solute molecules (e.g. sucrose) — for example, cellophane or an animal bladder membrane.

(ii) **Osmosis:**

The phenomenon where solvent molecules move through a semipermeable membrane into the solution side when a solvent and pure water (or solutions of different concentrations) are separated by the membrane.

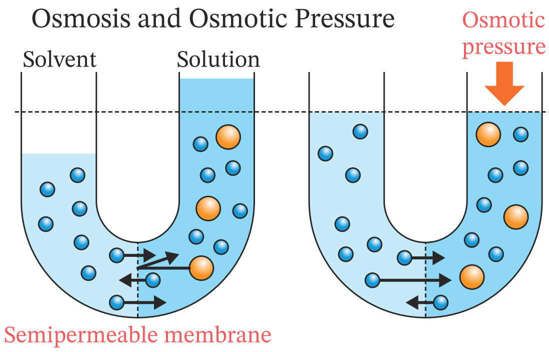


Figure 1.10.1 Osmosis and Osmotic Pressure

structure of science

Key Concepts:

- Semipermeable membrane
- Osmosis
- Osmotic pressure
- Reverse osmosis
- Concentration ·
- Solvent and solute.

Scientific law:

van't Hoff's law for dilute solutions ($\pi = cRT$).

Scientific Models

No scientist can directly observe billions of water molecules moving through a membrane at the same time. Instead, scientists use diagrams, models, and mathematical relationships to explain and predict what is happening.

Mechanism: Why does it occur?

Water molecules are constantly moving in both directions through the pores of the semipermeable membrane.

However, large "solute molecules" that cannot pass through the pores are present on the solution side, physically obstructing the return of water molecules to the pure water side. As a result, the number of water molecules moving from the "pure water → solution" side exceeds the number moving in the opposite direction, leading to a net movement of water into the solution.

(iii) **Osmotic pressure:**

- When osmosis occurs in a U-tube, the liquid level on the solution side rises. Osmosis stops when the pressure exerted by the difference in the height of the liquid levels (the hydrostatic pressure of the water column) balances the osmotic pressure.
- Therefore, osmotic pressure is equal to **the pressure that must be applied to the solution side to prevent osmosis (i.e., to maintain the liquid levels on both sides at the same height).**

🌋 Natural Phenomenon

Plant cells contain water that generates internal pressure against the cell walls, When water is lost, the cells become less rigid and the plant begins to wilt. Osmosis plays a crucial role in maintaining plant structure and survival.

⭐ Nature of Science

Scientific explanations are accepted because they are supported by evidence.

However, if new evidence emerges, scientists may revise or improve existing models.

(iv) van't Hoff's Law:

For dilute solutions, the osmotic pressure π [Pa] is proportional to the molar concentration c [mol/L] of the solution and the absolute temperature T [K], regardless of the type of solvent or solute.

$$\pi = cRT$$

(R: Molar gas constant. Since $c = \frac{n}{V}$, the relationship can also be expressed as $\pi V = nRT$. This equation has the same form as the ideal gas equation of state.)

📦 Equation box

Quantity	Relation
Van't Hoff's law	$\pi = cRT$
Equivalent form	$\pi V = nRT$ (R = molar gas constant)

Life Application

Reverse Osmosis-The Technology to Convert Saltwater into Fresh Water, What happens if pressure greater than the osmotic pressure is forcibly applied to the solution side?

Astonishingly, water molecules in the solution (the concentrated side) are forced through the semipermeable membrane and pushed out to the pure water (the dilute side), contrary to the natural direction of osmosis. This phenomenon is called reverse osmosis.

Reverse osmosis technology plays a vital role in supporting our daily lives, most notably in seawater desalination. By applying high pressure to seawater and passing it through a special semipermeable membrane (known as an RO membrane), solutes like salt are blocked, and only potable water (fresh water) is extracted. This chemical principle is applied in countries facing water scarcity, in water purification systems during disasters, and even for water recycling on the International Space Station (ISS).

An understanding of the invisible, microscopic phenomenon of water movement is connected to magnificent technology that solves global challenges.

Warm Up

Answer the following questions.

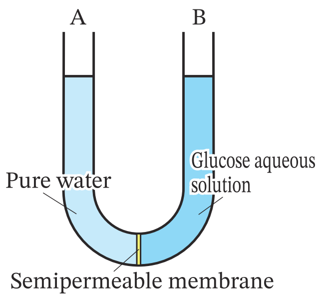
A As shown in the figure, the center of a U-tube is partitioned by a semipermeable membrane. Pure water is placed in A, and a glucose aqueous solution is placed in B at the same time so that both liquid levels are at the same height, and left at 27°C.

Assume the relative molecular mass of glucose is 180 and the gas constant is 8.3×10^3 Pa·L/(K·mol).

(a) After a while, which liquid level rises, A or B ?

(b) In the case of (a), write the name of the substance that moves through the semipermeable membrane.

(c) This glucose aqueous solution is made by dissolving 1.2 g of glucose $C_6H_{12}O_6$ to make 200 mL. What osmotic pressure (in Pa) must be applied to stop the level rising?



Explanation

(a) Since pure water permeates to the glucose aqueous solution side by osmosis, the liquid level on the glucose aqueous solution side rises. Therefore, **B**.

(b) Water

(c) When calculating the osmotic pressure, van't Hoff's law is used. Let the osmotic pressure be Π [Pa], from van't Hoff's law ΠV = nRT:

$$\Pi\ [Pa] \times 0.200\ L = \frac{1.2\ g}{180\ g/mol} \times 8.3\times10^3\ Pa\cdot L/(K\cdot mol) \times (27+273)\ K$$

Convert the unit of volume to [L] and the unit of temperature to [K].

$$\Pi = \frac{\cancelto{6\ 2}{1.2}\ g \times 8.3\times10^3\ Pa\cdot L/(K\cdot mol) \times \cancelto{5}{300}\ K}{\cancelto{3\ 1}{180}\ g/mol \times \cancelto{1}{0.200}\ L}$$

$= 8.3\times10^4$ Pa

Think Critically

A student says: «The water moved because the glucose molecules pulled it through the membrane».

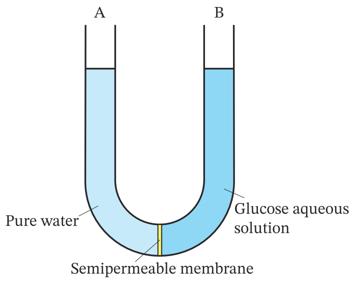
Evaluate the Claim

- Is this explanation fully accurate? ....
- What evidence supports or challenges it? ....
- How would you improve the explanation? ....

Try

. Read the following text and answer the questions below.

As shown in the figure, the center of a U-tube is partitioned by a semipermeable membrane. Pure water is placed in A , and a glucose aqueous solution is placed in B at the same time so that both liquid levels are at the same height, and left at 27°C. From the ((i) A / B) side to the ((ii) A / B) side, molecules of ((iii) water / glucose) pass through the semipermeable membrane, and the liquid level in ((iv) A / B) rises.



(a) Choose the appropriate letters or words for (i) to (iv) in the text from the brackets.

(b) This glucose aqueous solution is made by dissolving 1.8 g of glucose $C_6H_{12}O_6$ in water to make 300 mL. What is the pressure (osmotic pressure) in Pa applied to prevent the liquid level of this aqueous solution from rising? Calculate the osmotic pressure to two significant figures. Assume the gas constant R = 8.3×10^3 Pa·L/(K·mol). ($C_6H_{12}O_6$ = 180)

Answer: ___________________________________

Exercise

. As shown in the figure, the center of a U-tube is partitioned by a semipermeable membrane. A glucose aqueous solution is placed in A, and pure water is placed in B at the same time so that both liquid levels are at the same height, and left at 27°C. Answer the following questions.

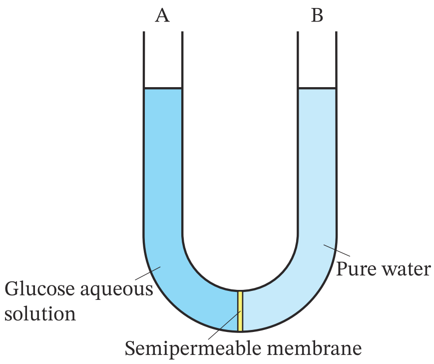
(a) Which liquid level rises, A or B?

____

(b) This glucose aqueous solution is made by dissolving 0.60 g of glucose $C_6H_{12}O_6$ in water to make 200 mL.

What is the pressure (osmotic pressure) in Pa applied to prevent the liquid level of this aqueous solution from rising? Calculate the osmotic pressure to two significant figures. Assume the gas constant $R = 8.3 \times 10^3$ Pa·L/(K·mol). ($C = 12$, $H = 1.0$, $O = 16$)

____



Challenge Your Thinking

A plant begins to wilt after being watered with a concentrated fertilizer solution.

Using your understanding of osmosis:

- Predict the direction of water movement. ....
- Explain why the plant loses rigidity. ....
- Suggest a solution to help the plant recover. ....

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

1-11 — Colloids

Engaging Introduction

You shine a laser beam through two containers — one holds pure water, the other holds diluted milk.

What is your observation?

____

What is your conclusion?

____

In this lesson we will investigate the reason behind this, understand the meaning of colloids, and study their main properties, including hydrophilic colloids and salting out, and hydrophobic colloids and coagulation.

❓ Guiding question: How can tiny particles dispersed in a medium influence the behaviour of light, motion, and matter?

Point!

1. Colloids

(a) **Colloidal particles:** particles about 10^{-9} m to 10^{-7} m in diameter — larger than ions and molecules; they cannot pass through a semipermeable membrane (e.g. cellophane) but can pass through filter paper.

(b) **Colloid:** a state in which colloidal particles are uniformly dispersed in another substance (or the substance itself).

(c) **Colloidal solution (sol):** a solution in which colloidal particles are dispersed in a liquid.

(d) **Gel:** a sol that has lost its fluidity. Example: hot milk with starch (sol) → chilled mahallabiya (gel).

Experimental Investigation

Materials

• Water • Milk • Salt solution • Starch solution • Laser pointer

Procedure

Shine the laser through each sample in a darkened room and observe whether the light path becomes visible.

What is your observation?..

What is your conclusion?..

structure of science

Key Fact :

The unique properties of colloids arise from the intermediate size and charge of their particles.

Key Concepts:

- Colloidal particles
- Sol (colloidal solution)
- Hydrophilic colloid
- Hydrophobic colloid
- Protective colloid.

2. Properties of colloids

(a) **Tyndall effect:** The phenomenon where colloidal particles scatter light, making the path of the light visible.

(b) **Brownian motion:** The irregular motion of colloidal particles caused by the collision of particles of the dispersion medium with the colloidal particles.

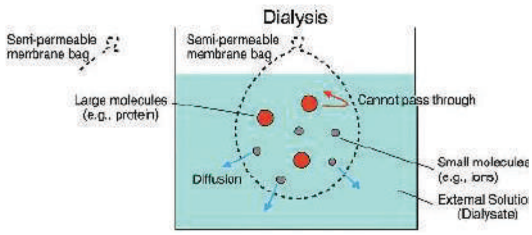


Figure 1.11.1 Dialysis Property

(c) **Dialysis:** The operation of separating and a colloidal solution using a semipermeable membrane, etc.(Because the colloidal particles are larger than the pores in the membrane).

(d) **Electrophoresis:** The phenomenon where charged colloidal particles move to one of the electrodes.

- Positively charged particles move to the cathode .
- Negatively charged particles move to the anode.

(e) **Hydrophilic colloids and salting out**

- **Hydrophilic colloid:** A colloid with a high affinity for water. (A large number of water molecules hydrate and disperse).
- **Salting out:** The phenomenon where precipitation occurs when a large amount of electrolyte is added to a hydrophilic colloid.

(f) **Hydrophobic colloids and coagulation**

- **Hydrophobic colloid:** A colloid with a low affinity for water. (Dispersed by the repulsive force between charges of the same sign). Example: Clay, sulfur, etc.
- **Coagulation:** The phenomenon where precipitation occurs when a small amount of electrolyte is added to a hydrophobic colloid.

An ion of the opposite sign to the charge of the colloidal particles and a higher valency is more likely to coagulate the colloid.

(g) **Protective colloid:** A hydrophilic colloid that makes it difficult for hydrophobic colloidal particles to coagulate when added to a hydrophobic colloid.

Warm Up

Answer the following questions.

(a) Read the text and answer.

A mixture in which particles of about 10^{-9} to 10^{-7} m are uniformly dispersed in a liquid is called a (A). When an aqueous solution of iron(III) chloride is added to boiling water, a (A) of iron(III) hydroxide forms. Placing this mixture in a cellophane bag immersed in distilled water removes the impurities and purifies it — this operation is called (B). When a small amount of electrolyte is added to a portion of the (A) and left to stand, precipitation occurs; this is called (C). Because this happens easily, the iron(III) hydroxide colloid is classified as a (D) colloid.

(i) Answer the words that correspond to A to D in the text.

(ii) The iron(III) hydroxide colloid is positively charged. Regarding the underlined part, which of the following aqueous electrolyte solutions (a to d) of the same molarity causes

Misconception Alert

«All mixtures that appear uniform are true solutions.»

Reality:

Some mixtures appear uniform but are actually colloids containing dispersed particles large enough to scatter light.

Interdisciplinary Connection

Hydrophobic colloids and coagulation

Colloids connect chemistry with:

- Biology (blood, proteins, cell membranes)
- Physics (light scattering, particle motion)
- Engineering (water treatment, pollution control)
- Medicine (dialysis).

precipitation in the smallest amount?

a. NaCl　　b. $Ca(NO_3)_2$　　c. K_2SO_4　　d. $CaCl_2$

(b) The following sentences describe characteristic phenomena exhibited by colloids. Identify the term closely related to each of the following phenomena.

(i) The path of light from car headlights can be seen from the side at night.

(ii) Soot can be removed by applying a high DC voltage inside a factory chimney.

(iii) Colloidal particles move irregularly when they collide with the surrounding solvent molecules.

Explanation

(a) (i) **A:** colloid (sol)

B: Small molecules and ions in the colloidal solution of iron(III) hydroxide pass through the cellophane membrane into the water. On the other hand, the colloidal particles of iron(III) hydroxide cannot pass through the cellophane membrane and remain in the bag, so the purity increases. The operation of separating a colloidal solution in this way is called dialysis. Dialysis.

C: The phenomenon where precipitation occurs when a small amount of electrolyte is added is coagulation.

Coagulation.

D: Hydrophobic

(ii) As it is a positively charged colloid, the electrolyte containing an anion with the highest ionic charge is the most

effective. Among **a** to **d**, the sulfate ion SO_4^{2-} has the largest valency. Therefore, **c**.

(b) (i) Colloidal particles in the air scatter light, making the path of light visible. Tyndall effect.

(ii) Colloidal particles of soot are attracted to the electrodes and are no longer emitted from the chimney.

Electrophoresis.

(iii) Brownian motion

Try

Choose the correct answer:

1. A beam of light becomes visible when it passes through a colloidal solution. Which property is being observed?

A. Brownian motion

B. Electrophoresis

C. Dialysis

D. Tyndall effect

Answer: ___________________________________

2. A colloidal solution is placed in an electric field and its particles move toward the anode. What can be concluded?

A. The particles are positively charged.

B. The particles are neutral.

Global Context

explanation

Colloidal chemistry plays an important role in Water purification, Air pollution control, Food production, Pharmaceutical industries.

Understanding colloidal systems helps scientists develop sustainable technologies for modern societies.

C. The particles are negatively charged.

D. The particles contain no water.

Answer: ____________________________________

3. Among the following descriptions about colloids from A to E, choose all that are correct.

A. Because colloidal particle diameter is about 10^{-3} m, they cannot pass through filter paper.

B. The absorption of light by a colloid is the Tyndall effect.

C. Brownian motion occurs because solvent molecules in thermal motion collide irregularly with the colloidal particles.

D. Adding a large amount of electrolyte to egg-white solution causes no precipitation.

E. A colloid that has lost its fluidity, like chilled mahallabiya, is called a sol.

Answer: ____________________________________

Exercise

Answer on the Following Questions

1. When a small amount of electrolyte is added to a colloidal solution of arsenic sulfide, colloidal particles gather and precipitate. Also, when two electrodes are immersed in this colloidal solution and (b)a DC voltage is applied, the colloidal particles gather on the anode side, so the color near the anode becomes darker.

(a) What are the phenomena of the underlined parts (a) and (b) called, respectively?

(b) Do the colloidal particles of arsenic sulfide carry a positive or negative charge?

(c) When the following aqueous electrolyte solutions from A to E are added to the colloidal solution, which one causes the coagulation described in (a) at the lowest molarity?

A. NaCl B. $MgCl_2$ C. $Al(NO_3)_3$ D. Na_2SO_4 E. KNO_3

Answer: ____________________________________

2. Match each of the following descriptions (i) to (v) with the correct phenomenon or operation (A to E).

(i) Strong light on a starch solution makes the light path bright. ·

(ii) When a DC voltage is applied to an iron

(III) hydroxide colloidal solution, the colloidal particles move to the cathode side.

(iii) Under an ultramicroscope, colloidal particles move irregularly.

(iv) A large amount of electrolyte added to starch-milk or gelatin causes precipitation.

(v) A small amount of electrolyte added to a sulfur colloid causes precipitation.

A. Salting out B. Coagulation C. Tyndall effect

D. Brownian motion E. Electrophoresis

Answer: ____________________________________

Challenge your thinking

A community suffers from muddy drinking water containing suspended colloidal particles.

Using concepts from this lesson to answer on this questions:

- Which colloidal properties could help remove these particles?............................
- Would filtration alone be sufficient?..
- How might coagulation improve purification?...

Support your proposal using scientific reasoning.

Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

1-12 — Structure of Metallic Crystals

Engaging Introduction

Look around you.

Steel bridges support enormous loads.

Airplanes made from metal travel thousands of kilometers.

Yet metals are made of tiny atoms separated by empty space.

In this lesson, we will explore how the three-dimensional arrangement of atoms inside metals determines their properties and performance.

> ? **Guiding question:** How does the arrangement of atoms inside a metal influence its properties and applications?

Point!

1. Crystals and amorphous solids

(a) **Crystal:** A solid in which the particles are arranged in a regular pattern.

(b) **Crystal lattice:** The regular arrangement structure of the particles in a crystal.

(c) **Unit cell:** The smallest repeating unit in a crystal lattice.

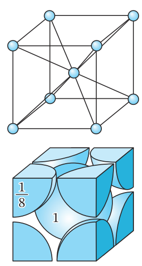
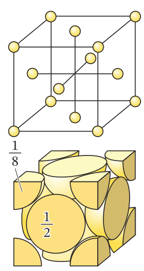
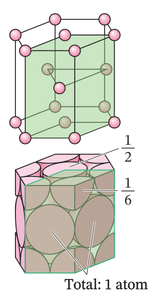
(d) **Coordination number:** The number of other particles present closest to a given particle in a crystal.

(e) **Amorphous solid:** A solid in which the arrangement of particles is not regular.

2. Structure of metallic crystals

(a) **Metallic crystal:** A crystal of many metal atoms held together by metallic bonds.

(b) **Structures of metallic crystals** — body-centred cubic (BCC), face-centred cubic (FCC), and hexagonal close-packed (HCP).

Name of structure	Body-centred cubic (BCC) lattice	Face-centred cubic (FCC) lattice	Hexagonal close-packed (HCP) structure
Unit cell			
Number of atoms in a unit cell	$\frac{1}{8}\times 8+1=$ 2	$\frac{1}{8}\times 8+\frac{1}{2}\times 6=$ 4	$\left(\frac{1}{6}\times 12+\frac{1}{2}\times 2+3\right)\times\frac{1}{3}=$ 2
Coordination number	8	12	12
examples	Iron	Copper	Titanium

structure of science

Key Fact :

The properties of a crystalline solid depend on the regular arrangement of particles within its crystal lattice.

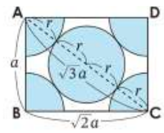
Key Concepts:

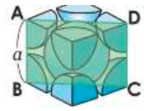
- Crystal
- Crystal lattice
- Coordination number
- Amorphous solid
- Metallic crystal
- Atomic radius ·
- Packing efficiency.

Structure	Atoms per unit cell	Coordination number
BCC	(1/8 × 8) + 1 = 2	8
FCC	(1/8 × 8) + (1/2 × 6) = 4	12
HCP	2	12

(c) Relationship between the edge length of a unit cell and the atomic radius Let the edge length of the unit cell be a and the atomic radius be r.

(i) Body-centred cubic lattice: In the plane ABCD of the figure on the right,



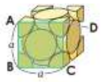


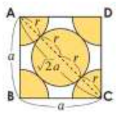
since AB = a, BC = $\sqrt{2}a$, we obtain, AC = $\sqrt{3}a$.

It can be calculated using Pythagoras' theorem.

Also, since AC = 4r, 4r = $\sqrt{3}a \Rightarrow r = \dfrac{\sqrt{3}}{4}a$

(ii)Face-centred cubic lattice: In the plane ABCD of the figure on the right, since AB = BC =a, AC = $\sqrt{2}a$.





Also, since AC = 4r, 4r = $\sqrt{2}a \Rightarrow r = \dfrac{\sqrt{2}}{4}a$.

(d) **Packing efficiency:** the percentage of the unit cell's volume occupied by atoms = **(volume of atoms in the unit cell) ÷ (volume of the unit cell)**.

Equation box — metallic structures at a glance

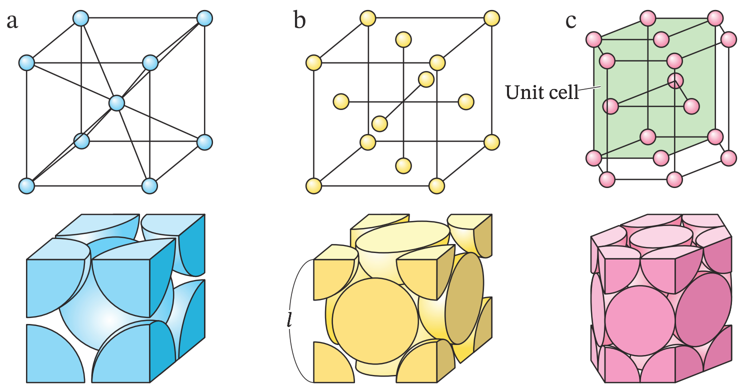
Structure	Atoms/cell	Coordination no.	r vs a
BCC	2	8	r = ($\sqrt{3}$ / 4) a
FCC	4	12	r = ($\sqrt{2}$ / 4) a
HCP	2	12	—

Life Application

Electrical Wiring & Devices: The face-centered cubic structure of copper allows electrons to flow freely. This provides the high electrical and thermal conductivity essential for household wiring, printed circuit boards (PCBs), and modern computing.

Warm Up

Answer the following questions.



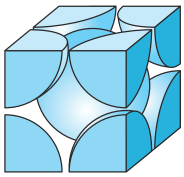
(a) The figures show the structures of metallic crystals (a, b, c).

(i) What are the crystal structures of a and c called?

(ii) How many atoms are contained in the unit cell of a?

(iii) In b, letting the atomic radius be r and the edge length of the unit cell be l, express r in terms of l.

(b) The figure on the right represents the crystal lattice of iron. Letting the radius of an iron atom be r (cm) and the edge length of the crystal lattice in the figure be a (cm), answer the following questions



(i) Express r in terms of a.

(ii) Find the packing efficiency of iron atoms. You may leave square roots in your expression.

(iii) Find the density of iron atoms (g/cm3). Assume the Avogadro constant is NA (/mol). (Fe = 56)

Explanation

(a) (i) a: Body-centred cubic lattice / c: Hexagonal close-packed structure (ii) 2

(iii) The crystal structure of b is a face-centred cubic lattice. Therefore, $r = \frac{\sqrt{2}}{4}l.$

(b) (i) From the figure, the crystal lattice of iron is a body-centred cubic lattice. Therefore, $r = \frac{\sqrt{3}}{4}a.$

(ii) Packing efficiency = $\frac{\text{Volume of atoms in the unit cell}}{\text{Volume of the unit cell}}$

The number of atoms per unit cell is 2, and their volume is:

$$\frac{4\pi r^3}{3} \times 2 = \frac{4\pi \times \left(\frac{\sqrt{3}}{4}a\right)^3}{3} \times 2 = \frac{\sqrt{3}\pi a^3}{8} [\text{cm}^3]$$

Since the volume of the unit cell is a^3 [cm^3], the packing efficiency is:

$$\frac{\frac{\sqrt{3}\pi a^3}{8}}{a^3} = \frac{\sqrt{3}\pi}{8}$$

(iii) For the density of atoms, determine the density of the unit cell.

Density = $\frac{\text{Mass of the unit cell}}{\text{Volume of the unit cell}}$

First, since the molar mass of iron Fe is 56 g/mol, the mass of 2 iron atoms in the unit cell is:

$$\frac{56\text{ g/mol} \times 2}{N_A[\text{/mol}]} = \frac{112}{N_A} [\text{g}]$$

Since the volume of the unit cell is a^3 [cm^3], the density is:

$$\frac{\frac{112}{N_A}[\text{g}]}{a^3[\text{cm}^3]} = \frac{112}{a^3 N_A} [\text{g/cm}^3]$$. Answer: $\frac{112}{a^3 N_A} [\text{g/cm}^3]$

Design Models by Home Materials

Body-Centered Cubic (BCC)

Materials

- 9 clay balls
- Toothpicks

Procedure

1. Construct a cube using 8 clay balls at the corners.
2. Connect corners using toothpicks.
3. Place one clay ball in the center of the cube.
4. Observe the arrangement from different directions.

Investigation

- Count corner atoms.

Answer: ________________________________

- Identify the central atom.

Answer: ________________________________

- Determine the coordination number.

Answer: ________________________________

You can design model for Face-Centered Cubic (FCC) and Hexagonal Close-Packed (HCP), by the same procedure .

Try

Choose the correct answer:

1. A metallic crystal has a coordination number of 8. Which structure is most likely present?

A. FCC B. HCP

C. BCC D. Amorphous structure

Answer: ________________________________

2. Both FCC and HCP structures have a coordination number of:

A. 6 B. 8

C. 10 D. 12

Answer on the following questions:

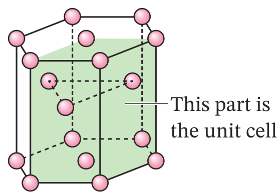
3. The figure shows the crystal structure of a certain metal. Answer the following questions.

(a) What is the crystal lattice of this metal called?

____

(b) What is the total number of atoms contained in a unit cell?

____



(c) What is the coordination number of an atom in this lattice?

____

(d) A solid with particles arranged regularly, like a metallic solid, is called a crystal. On the other hand, what is a solid with an irregular arrangement that is not crystallized called?

____

Exercise

Answer on the Following Questions

(a) Name the three types of crystal lattice found in metallic crystals.

____

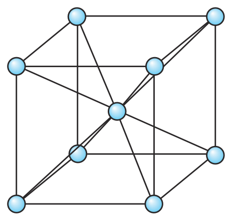
(b) Of the three types named in (a), which has the largest number of atoms in a unit cell?

____

(c) Of the three types named in (a), which has the smallest coordination number? Also, write the coordination number of that lattice

____

2. The figure on the right is the crystal structure of iron. Answer the following questions.



(a) In the iron crystal, what is the total number of atoms contained in a unit cell?

____

(b) Assuming the edge length of the iron unit cell is a [cm], express the atomic radius of iron in terms of a. You may leave square roots in your expression.

____

(c) Express the packing efficiency of the iron crystal lattice in terms of π.

____

(d) Assuming the density of iron is d (g/cm^3), express the mass of one iron atom in terms of a and d.

____

(e) Assuming the Avogadro constant is N_A (/mol), express the relative atomic mass of iron in terms of a, d, and N_A

____

Think Critically

You are a materials engineer tasked with designing a metal for a lightweight electric vehicle.

Which crystal structure would you prefer?

- BCC
- FCC
- HCP

Answer: ______________________________________

Justify your choice using:

- Packing efficiency
- Coordination number

Answer: ______________________________________

1-13 — Structure of Ionic Crystals

✨ Engaging Introduction

Every grain of table salt contains billions of sodium and chloride ions. These ions carry opposite charges and are held together by powerful electrostatic attractions, so salt crystals grow into regular geometric shapes visible to the naked eye. In this lesson we explore how the arrangement of ions within a crystal determines its structure, stability, and physical properties.

❓ **Guiding question:** How does the arrangement of ions within a crystal determine its structure, stability, and physical properties?

Exploration

Using colored balls:
• Blue = Na^+
• Green = Cl^-
Arrange ions to satisfy the following condition:
Every ion should be surrounded by as many oppositely charged ions as possible.
Explore
• What pattern emerges?
• Why do ions not arrange randomly?

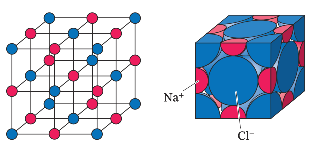
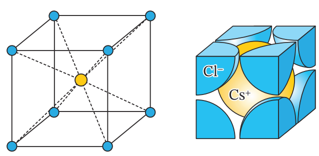
📌 Point!

Structure of ionic crystals

(a) **Ionic crystal:** A crystal formed by cations and anions bound together by electrostatic forces (ionic bonds).

(b) **Coordination number:** the number of oppositely charged ions immediately surrounding a given ion in an ionic crystal.

(c) **Structural examples** — sodium chloride (NaCl) and caesium chloride (CsCl).

Ionic crystal	Sodium chloride NaCl	Caesium chloride CsCl
Unit Cell		
Number of ions in a unit cell	Na^+: $\frac{1}{4}\times12+1=$ 4 Cl^-: $\frac{1}{8}\times8+\frac{1}{2}\times6=$ 4	Cs^+: 1 Cl^-: $\frac{1}{8}\times8=$ 1
Coordination number	Na^+: 6 Cl^-: 6	Cs^+: 8 Cl^-: 8

⭐ structure of science

Key Fact :

The properties of ionic crystals depend on the regular three-dimensional arrangement of oppositely charged ions within the crystal lattice.

Key Concepts:

- Crystal
- Crystal lattice
- Cation
- Anion
- Ionic bond
- Unit cell
- Coordination
- number
- Ionic radius

Equation box

Quantity	Relation
Density	formula units per cell × formula mass) ÷ (N_A × a^3
NaCl edge contact	a = 2·r(Na^+) + 2·r(Cl^-)
CsCl body diagonal	√3·a = 2·r(Cs^+) + 2·r(Cl^-)

Real-World Application

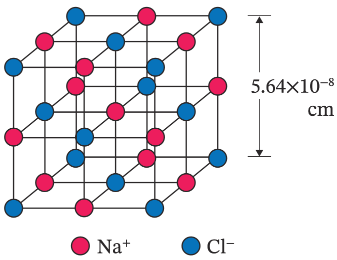
Salt in Food Preservation Salt makes water less available to microorganisms, which helps preserve food and prevent spoilage. Understanding ionic compounds and how they interact with water is essential in food science and preservation technologies.

Warm Up

The figure shows the NaCl crystal lattice, edge length 5.64×10^{-8} cm. (N_A = 6.0×10^{23} /mol; $5.64^3 = 179$; Na = 23, Cl = 35.5.)

(a) If the ionic radius of Na^+ is 1.16×10^{-8} cm, find the ionic radius r [cm] of Cl^- (three significant figures).

(b) Find the density of the NaCl crystal in g/cm^3 (to two decimal places).

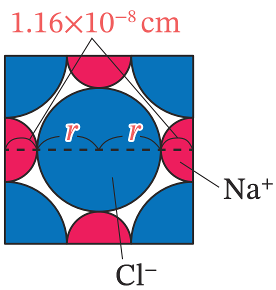


Explanation

(a) Along the edge where Na^+ and Cl^- are in contact, the edge length of the unit cell can be expressed as $1.16{\times}10^{-8}\text{ cm} \times 2 + r\,[\text{cm}]{\times}2$.

Therefore, $1.16{\times}10^{-8}\text{ cm} \times 2 + r\,[\text{cm}]{\times}2 = 5.64{\times}10^{-8}$ cm.

$r = 1.66{\times}10^{-8}$ cm



(b) Density is defined as: Density $= \dfrac{\text{Mass of unit cell}}{\text{Volume of unit cell}}$.

A unit cell of sodium chloride contains four Na^+ ions and four Cl^- ions, corresponding to four formula units of NaCl.

The mass of 4 NaCl units is: $\dfrac{58.5\text{ g/mol}}{6.0{\times}10^{23}\text{/mol}}{\times}4 = \dfrac{39}{10^{23}}$ g.

Also, since the volume of the unit cell is $(5.64{\times}10^{-8})^3\text{cm}^3$, the density of the crystal is:

$$\frac{\dfrac{39}{10^{23}}\text{g}}{(5.64{\times}10^{-8})^3\text{cm}^3} = 2.17\text{ g/cm}^3$$

Try

1. What is the primary force holding an ionic crystal together?

A. Hydrogen bonding B. Metallic bonding C. Electrostatic attraction between oppositely charged ions D. Van der Waals forces

Answer: ____

2. What is meant by the coordination number of an ion in an ionic crystal?

A. The number of electrons in the ion. B. The number of ions in the unit cell. C. The number

Nature of Science

Scientists developed crystal models using evidence from X-ray diffraction experiments, mathematical analysis, and repeated experimental observations.

of charged ions surrounding that ion. D. The number of bonds formed by the ion.

Answer: ___________________________________

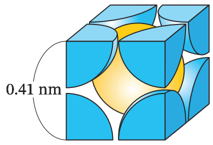
3. The figure is the CsCl unit cell, with Cs^+ and Cl^- in contact. (Two significant figures; edge = 0.41 nm; $\sqrt{3} = 1.7$; N_A = 6.0×10^{23} /mol; Cl = 35.5, Cs = 133.)

(a) If the radius of Cl^- is 0.17 nm, find the radius of Cs^+ in nm.

(b) Find the volume of 1 mol of CsCl crystal in cm^3 (1.0 nm = 1.0×10^{-7} cm; $4.1^3 = 69$).

(c) Find the density of the CsCl crystal in g/cm^3.

Answer: ___________________________________



Exercise

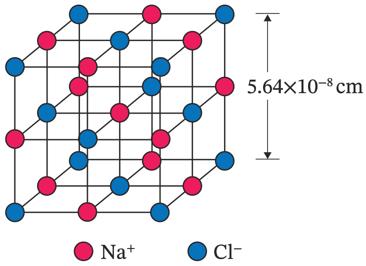
1. The edge of the unit cell of a NaCl crystal is 5.64×10^{-8} cm. Answer the following questions. Assume the Avogadro constant is 6.0×10^{23} /mol.

(a) What are the numbers of Na^+ and Cl^- in a unit cell, respectively?

(b) What is the total mass of NaCl in a unit cell in grams?

(c) What is the volume of the unit cell in cm^3? Assume $5.64^3 = 1.8\times10^2$.

(d) Calculate the density of the crystal in g/cm^3 to two decimal places.



Reflect and Grow

Complete:

✓ One idea that changed my thinking

....

✓ One real-life application I discovered

....

✓ One question I still want to investigate

....

Unit question

Read this case study, then Answer on the following questions

A humanitarian organization is designing portable cooling boxes to transport vaccines to remote communities where electricity is unavailable. The engineering team is considering three substances as cooling materials.

Data Table

Substance	Molar Mass (g/mol)	Heat of Fusion (kJ/mol)	Heat of Vaporization (kJ/mol)	Dominant Inter-molecular Force	Boiling Point (°C)
Water (H_2O)	18	6.0	41	Hydrogen bonding	100
Hydrogen Sulfide (H_2S)	34	2.4	18	Dipole–dipole / Van der Waals	-60
Hydrogen Bromide (HBr)	81	2.2	17	Dipole–dipole / Van der Waals	-67

The cooling box must maintain a low temperature for the longest possible time while remaining safe and practical for transport.

Part A — Data Analysis (Short Response)

Analyze the data provided and identify which substance would absorb the greatest amount of energy during a phase change from liquid to gas.

....

Support your answer using quantitative evidence from the table.

....

Part B — Scientific Explanation (Structured Response)

The boiling point of water is significantly higher than that of H_2S, even though water has a lower molar mass.

Explain this observation using intermolecular forces and particle interactions.

....

Part C — Evaluation (Extended Response)

An engineer argues:

“A substance with the largest molar mass should always have the highest boiling point because heavier molecules require more energy to separate.”

Using evidence from the table and your understanding of intermolecular forces:

Evaluate the validity of this claim.

In your answer:

- discuss supporting evidence,
- identify limitations of the claim,
- compare molecular mass and intermolecular forces as factors affecting boiling point.

....

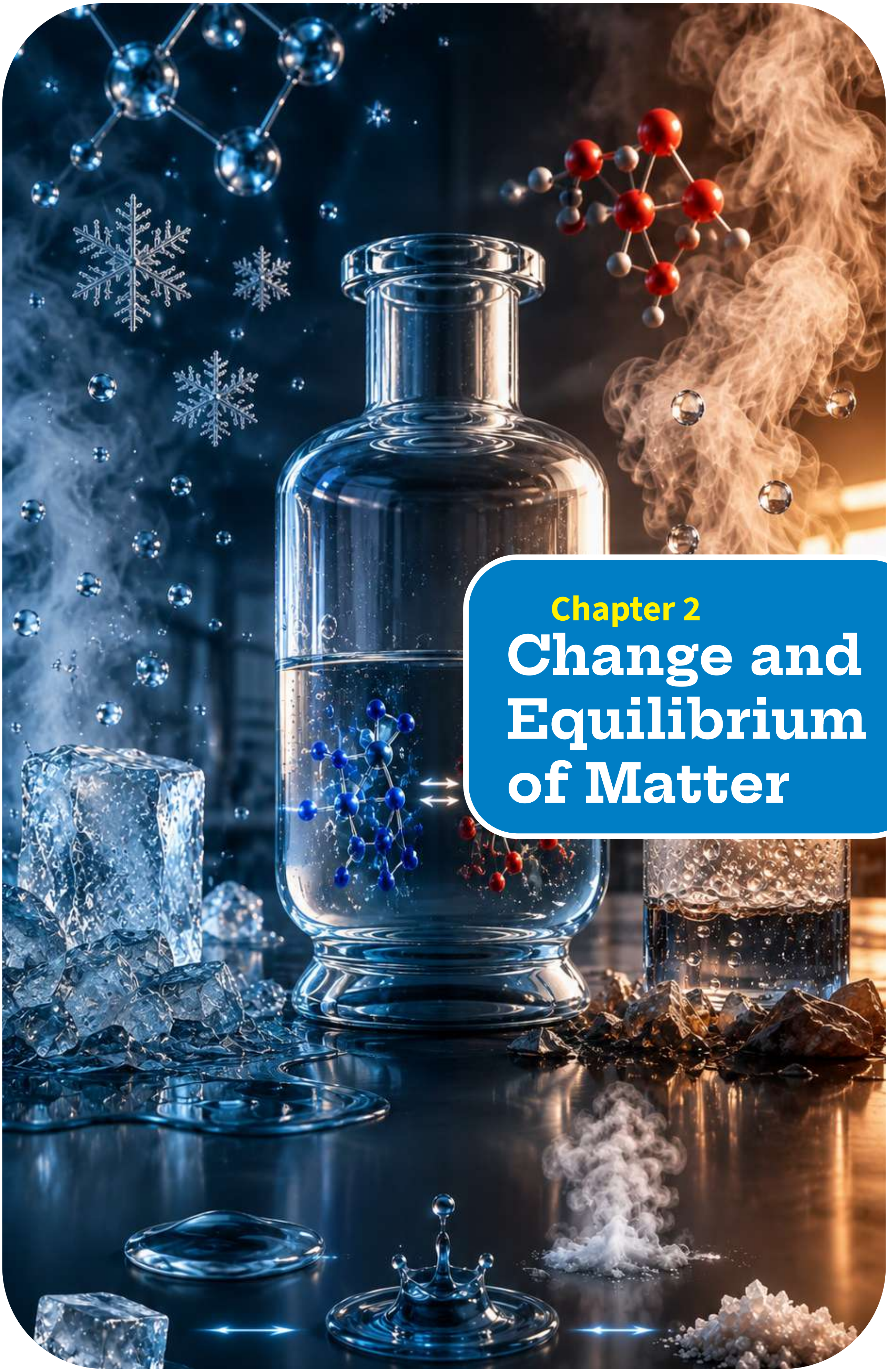
Part D — Essay Question (Extended Response)

To what extent is intermolecular force strength more important than molecular mass in determining the physical properties of substances?

Using evidence from the case study and examples from the unit, construct a well-reasoned scientific argument that:

- compares hydrogen bonding and Van der Waals forces,
- explains their effects on melting and boiling points,
- discusses the role of energy transfer during changes of state,
- reaches a justified conclusion supported by scientific evidence.

....



Chapter 2

Change and Equilibrium of Matter

2-1 — Chemical Reactions and Enthalpy Changes

✨ Engaging Introduction

Imagine holding an instant cold pack after a sports injury. As soon as you squeeze it, it suddenly becomes cold without ever being put in a refrigerator. In another situation, a campfire releases heat and light that keep people warm on a cold night. In this lesson we will explore how chemists track the energy changes during chemical reactions, and how these invisible energy transfers shape the world around us.

❓ **Guiding question:** How can scientists use energy changes in chemical reactions to explain and predict physical and chemical phenomena?

🔭 Exploration

Hot or Cold?

Consider the following situations:

• Burning natural gas in a kitchen stove.

• Dissolving ammonium nitrate in water inside a cold pack.

• Evaporation of sweat from the skin.

Without studying the lesson:

• Which processes release energy? ....

• Which processes absorb energy? ....

Record your initial ideas and revisit them after studying enthalpy changes.

....

📌 Point!

1. Chemical reactions and enthalpy

(a) Chemical reactions and heat transfer

(i) Exothermic reaction: A reaction that releases heat to the surroundings.

(ii) Endothermic reaction: A reaction that absorbs heat from the surroundings.

(b) __Enthalpy__ : A quantity representing the heat content (energy) of a substance.
Denoted by the symbol (H).
The unit is Joules (Symbol: J).

(c) __Enthalpy of reaction__ : The change in enthalpy during a chemical reaction.
Denoted by the symbol ΔH.

(d) Sign of enthalpy of reaction

The value of ΔH is __negative__ for exothermic reactions and __positive__ for endothermic reactions.

Exothermic reactions release heat
Endothermic reactions absorb heat

⭐ **structure of science**

Key Fact :

Exothermic reactions release heat to the surroundings.
- Endothermic reactions absorb heat from the surroundings. -
The enthalpy of combustion is always negative.

Key Concepts:

· Chemical reaction
· Heat transfer
· Enthalpy (H)
· Enthalpy change (ΔH)
· Exothermic reaction
· Endothermic reaction
· Enthalpy of combustion
· Enthalpy of formation
· Enthalpy of solution
· Enthalpy of neutralisation.

Investigating Energy Changes

Students observe two safe demonstrations:

Investigation A

Dissolve calcium chloride in water.

Investigation B

Dissolve ammonium nitrate in water.

Students:

- Measure temperature before and after mixing.
- What is your observations? ....
- Classify each process as exothermic or endothermic. ....
- Infer the direction of energy transfer. ....

Inquiry Question

What evidence allows scientists to conclude that energy has entered or left a system?

....

2. Expressing the enthalpy of reaction

1. Write the chemical equation.
2. Write the physical state of the substance after the chemical formula.

   Note: If allotropes exist, write the name of substance .

   Example: C (graphite)
3. Make the coefficient of the substance of interest 1 .
4. Write ΔH after the reaction equation .

Example: The reaction where 1 mol of hydrogen H_2 burns to produce water H_2O
(Enthalpy of reaction is −286 kJ)

1. $2H_2 + O_2 \longrightarrow 2H_2O$
2. $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$
3. $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)}$
4. $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)} \quad \Delta H = -286\ kJ$

Think Critically

A company advertises:

«Our new fuel releases twice as much heat as ordinary fuel, therefore it is always better for the environment.»

Claim–Evidence–Reasoning

Students evaluate:

- Is the claim scientifically valid? ....
- What evidence is missing? ....
- Does higher energy release always mean better sustainability?..........................

3. Types of enthalpy of reaction

(a) Enthalpy change during phase change: When a substance changes state, its enthalpy changes.The enthalpy of phase change is defined as the enthalpy change when 1mol of a substance undergoes a phase transition.

(b) **Enthalpy of combustion:** The enthalpy of reaction when 1mol of a substance undergoes complete combustion. It always takes a negative value. Because heat is released

(c) **Enthalpy of solution:** The enthalpy of reaction when 1 mol of a substance dissolves in a large amount of solvent.

- The dissolution of a substance in water is expressed using "aq", which represents a large amount of water.

Example: The enthalpy of solution of sodium hydroxide NaOH is −44.5 kJ/mol.

$$NaOH_{"s"} \longrightarrow NaOH_{"aq"} \quad \Delta H = -44.5\ kJ$$

(d) **Enthalpy of neutralisation:** The enthalpy of reaction when an acid and a base react to form 1 mol of water.

(e) **Enthalpy of formation**: The enthalpy of reaction when 1 mol of a compound is formed from its constituent elements in their standard states.

Types of reaction enthalpy at a glance

Type	Defined per…	Sign
Combustion	1 mol substance fully burned	always negative
Solution	1 mol dissolved in much solvent	±
Neutralisation	1 mol water formed (acid + base)	negative
Formation	1 mol compound from elements	±
Phase change	1 mol undergoing a transition	±

Life Application

Instant Hot and Cold Packs: These rely on specific enthalpy changes. An instant hot pack utilizes an exothermic reaction (like magnesium and water or a combination of iron, carbon, and a salt mixture when the pack is squeezed or crushed), releasing heat to your surroundings. A cold pack utilizes an endothermic reaction (like dissolving ammonium nitrate in water), absorbing heat and cooling your skin.

Warm Up

Answer the following questions.

(a) Explain what ΔH represents in the following chemical equation.

Identify all substances and types of enthalpy that could be described by this ΔH value.

$$C(graphite) + O_{2(g)} \longrightarrow CO_{2\,(g)}\ \Delta H = -394\ kJ.$$

(b) Express the following chemical changes or phase transitions as thermochemical equations.

(i) When 1 mol of ethane C_2H_6 undergoes complete combustion, 1562 kJ of heat is released.

Be Global Citizen

Understanding enthalpy changes helps scientists:

- Design cleaner fuels.
- Improve battery technologies.
- Develop efficient industrial processes.
- Reduce global energy consumption.

Many modern sustainability initiatives rely on controlling energy transfers in chemical reactions.

(ii) When 9.0 g of water evaporated, it absorbed 22 kJ of heat from the surroundings. (H = 1.0, O = 16).

Explanation

(a) The substance of interest in the chemical equation always has a coefficient of 1.

First, since 1 mol of graphite reacts with oxygen, it can be considered a reaction where graphite undergoes complete combustion. In this case, ΔH represents the enthalpy of combustion of graphite.

Also, it is a reaction where 1 mol of carbon dioxide is formed from carbon (graphite) and oxygen, so in this case, it represents the enthalpy of formation of carbon dioxide.

Based on the above, the value represents the enthalpy of combustion of graphite and the enthalpy of formation of carbon dioxide.

(b) Write according to the procedure.

(i) ❶ Write the chemical equation.

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

❷ Write the physical state of the substance after the chemical formula.

$$2C_2H_{6(g)} + 7O_{2(g)} \longrightarrow 4CO_{2(g)} + 6H_2O_{(l)}$$

❸ Since it represents the reaction of 1 mol of ethane, set the coefficient of C_2H_6 to 1.

$$C_2H_{6(g)} + \frac{7}{2}O_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(l)}$$

❹ Write ΔH after the reaction equation.

Noting that the enthalpy of combustion is negative, the enthalpy of combustion of 1 mol of ethane is −1562 kJ. Therefore,

$$C_2H_{6(g)} + \frac{7}{2}O_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(l)} \quad \Delta H = -1562 \text{ kJ}$$

(ii) ❶ ❷ Write the reaction equation.

$$H_2O_{(l)} \longrightarrow H_2O_{(g)}$$

❸ ❹ Since the coefficient of H_2O is 1, write ΔH after the reaction equation.

Note here that the enthalpy of reaction is expressed per 1 mol of the substance of interest.

The amount of substance for 9.0 g of water is $\frac{9.0 \text{ g}}{18 \text{ g/mol}} = 0.50$ mol

Therefore, since the heat absorbed when 0.50 mol of water evaporates is 22 kJ, the enthalpy of reaction when 1 mol of water evaporates is 44 kJ.

Thus, $H_2O_{(l)} \longrightarrow H_2O_{(g)} \quad \Delta H = 44 \text{ kJ}$

Data Interpretation Activity

Analyze Experimental Data

Substance Dissolved	Initial Temp (°C)	Final Temp (°C)
NaOH	25	34
NH_4NO_3	25	18

Think Like a Chemist

1. Which process is exothermic? ....

2. Which process is endothermic? ....

3. What evidence supports your conclusion?..

4. Which process has $\Delta H > 0$? ....

Try

Answer on the Following Questions.

1. State the specific type of reaction enthalpy represented by ΔH in the following chemical equation, and specify the substance it refers to.

(a) $HCl_{(aq)} + KOH_{(aq)} \longrightarrow KCl_{(aq)} + H_2O_{(l)} \quad \Delta H = -56\ kJ$

(b) $C_2H_{6(g)} + \frac{7}{2}O_{2(g)} \longrightarrow 2CO_{2(g)} + 3H_2O_{(l)} \quad \Delta H = -1562\ kJ$

2. Express the following changes as thermochemical equations. (H = 1.0, N = 14, O = 16)

(a) When 1 mol of methane CH_4 undergoes complete combustion producing carbon dioxide and water, 891 kJ of heat is released.

(b) When 1 mol of hydrogen H_2 burns to produce gaseous H_2O, 242 kJ of heat is released.

(c) When 1 mol of graphite is exposed to steam to produce carbon monoxide and hydrogen, 131 kJ of heat is absorbed.

(d) When 4.0 g of ammonium nitrate NH_4NO_3 is dissolved in a large amount of water, 1.3 kJ of heat is absorbed.

Answer: ______________________________________

3. Express the following phase transitions as thermochemical equations.

(a) The enthalpy of vapourisation of bromine is 31 kJ/mol.

(b) The enthalpy of condensation of ethanol C_2H_5OH is −39 kJ/mol.

Exercise

Choose the correct answer:

1. Why is the enthalpy of combustion always negative?

A. Products contain no energy.

B. Combustion absorbs heat.

C. Heat is released to the surroundings.

D. Oxygen is consumed.

Answer: ______________________________________

2. Which type of enthalpy change occurs when an acid reacts with a base to form water?

A. Enthalpy of formation

B. Enthalpy of combustion

C. Enthalpy of solution

D. Enthalpy of neutralization

Answer: ___________________________________

Answer on the following questions

3. State the specific type of reaction enthalpy represented by ΔH in the following chemical equation, and specify the substance it refers to.

(a) $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$ $\Delta H = -56$ kJ

(b) $CO_{(g)} + \frac{1}{2}O_{2(g)} \longrightarrow CO_{2(g)}$ $\Delta H = -283$ kJ

Answer: ___________________________________

4. Express each change as a thermochemical equation. (H = 1.0, O = 16, Na = 23)

(a) When 1 mol of diamond undergoes complete combustion, 395 kJ of heat is released.

(b) When 4.0 g of solid NaOH is dissolved in a large amount of water, 4.5 kJ of heat is released.

(c) When 1 mol of nitric oxide NO is formed from nitrogen N_2 and oxygen O_2, 90 kJ of heat is absorbed.

Answer: ___________________________________

5. Express each phase transition as a thermochemical equation.

(a) The enthalpy of vaporisation of ethanol C_2H_5OH is 39 kJ/mol.

(b) The enthalpy of sublimation of iodine is 62 kJ/mol.

Answer: ___________________________________

Challenge Your Thinking

A chemical reaction has:

$\Delta H = -250$ kJ

A student claims:

«The reaction absorbed 250 kJ of energy because the magnitude is large.»

Evaluate the claim.

- Is the reasoning correct? ....
- what sign is meant? ....
- How would the surroundings be affected? ....
- Use scientific evidence to justify your answer. ....

2-2 — Hess's Law

Engaging Introduction

In chemistry, reactions can follow different pathways — and, surprisingly, the total energy change stays the same regardless of the route. In this lesson we explore how chemists determine unknown energy changes using Hess›s law.

❓ Guiding question: How can scientists determine the energy change of a reaction that cannot be measured directly?

Exploration

Students examine three reactions: A → B · B → C · A → C. Investigate: - Can ΔH(A→C) be predicted from the other two reactions? .... - What evidence supports your conclusion? ....

Point!

Hess's law

(a) **Hess's law:** The law stating that if the initial and final states of a reaction are determined, the total enthalpy of reaction is constant, regardless of the reaction pathway.

(b) **Calculating the enthalpy of reaction using Hess's law:**

1. Let the required enthalpy of reaction be x [kJ/mol], and write the chemical equation for the target reaction.
2. Write all the chemical equations for the reactions whose enthalpies of reaction are given.
3. Combine the chemical equations written in 2 to form the chemical equation in 1.
4. Calculate the required enthalpy of reaction.

Example: When the enthalpy of combustion of graphite C is −394 kJ/mol and the enthalpy of combustion of carbon monoxide CO is −283 kJ/mol, find the enthalpy of formation of carbon monoxide.

1. Let the required enthalpy of formation be x [kJ/mol], and write the chemical equation representing the formation of carbon monoxide CO.

$$C(graphite) + \frac{1}{2}O_{2(g)} \longrightarrow CO_{(g)} \quad \Delta H = x\ [kJ]$$

Expressed so that the coefficient of CO becomes 1

2. Write the chemical equations for the combustion of C (graphite) and CO, for which the enthalpies of reaction are given.

$$C(graphite) + O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H = -394\ kJ \ ......(i)$$

$$CO_{(g)} + \frac{1}{2}O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H = -283\ kJ \ ......(ii)$$

3. Combine the chemical equations written in 2 to form the chemical equation in 1.

In the chemical equation of 1,
The coefficient of C is 1 and it is on the left side of the arrow, so add equation (i) in the same direction.
The coefficient of CO is 1 and it is on the right side of the arrow, so add equation (ii) in the reverse direction.

$$C(graphite) + O_2(g) \longrightarrow CO_2(g)$$

$$+)\ CO_2(g) \longrightarrow CO(g) + \frac{1}{2}O_2(g)$$

$$C(graphite) + \cancel{O_2}(g) + \cancel{CO_2(g)}$$

$$\longrightarrow \overset{\frac{1}{2}O_2}{\cancel{CO_2(g)}} + CO(g) + \cancel{\frac{1}{2}O_2(g)}$$

$$\Rightarrow C(graphite) + \frac{1}{2}O2(g) \longrightarrow CO(g)$$

structure of science

Key Fact :

The total enthalpy change of a reaction is the same regardless of the pathway taken between the reactants and products.

Key Concepts:

- Enthalpy change (ΔH)
- Hess's law
- Enthalpy of formation
- Enthalpy of combustion
- Bond enthalpy
- Thermochemical equations.

Scientific law

Hess's law — the total enthalpy change is independent of the reaction pathway.

When a reaction is reversed, consider the equation as subtracted .
Therefore, calculating (i) − (ii) yields the chemical equation of ❶ .

❹ Calculating the enthalpy of reaction in the same way as the transformation made in ❸

$$x = -394-(-283) \quad = -111 \ [kJ]$$

(Enthalpy of reaction of equation (i)) − (Enthalpy of reaction of equation (ii))

From the above, the enthalpy of formation of CO is −111 kJ/mol

(c) Bond enthalpy: The energy required to break covalent bonds within a molecule. Expressed as a value per 1 mol of bonds in the molecule.

• In a single molecule, if there are multiple bonds between atoms, the bond enthalpy represents the energy required to break all the bonds.

Example: The reaction converting methane CH4 into a carbon atom and hydrogen atoms

$$CH_{4(g)} \longrightarrow C_{(g)} + 4H_{(g)}$$

The bond enthalpy of this reaction is the energy to break 4 C–H bonds.

Key tools

Tool	Rule
Reverse a reaction	change the sign of ΔH (subtract it)
Multiply a reaction ×n	multiply ΔH by n
From bond enthalpies	ΔH = Σ(bonds broken) − Σ(bonds formed)
From formation enthalpies	ΔH = Σ ΔH_f(products) − Σ ΔH_f(reactants)

Warm Up

Answer the following questions.

(a) Calculate the enthalpy of formation [kJ/mol] of propane C_3H_8. Assume that the enthalpies of combustion of graphite, hydrogen, and propane are −394 kJ/mol, −286 kJ/mol, and −2219 kJ/mol, respectively.

(b) Using the following chemical equations, calculate the bond enthalpy [kJ/mol] of the O-H bond, and round your answer to the nearest integer.

$$H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(g)} \quad \Delta H = -242 \text{ kJ}$$

$$H_{2(g)} \longrightarrow 2H_{(g)} \quad \Delta H = 436 \text{ kJ}$$

$$O_{2(g)} \longrightarrow 2O_{(g)} \quad \Delta H = 498 \text{ kJ}$$

Explanation

(a) ❶ Write the chemical equation representing the formation of propane C_3H_8. Let x [kJ/mol] represent the enthalpy of formation of propane.

$$3C(\text{graphite}) + 4H_{2(g)} \longrightarrow C_3H_{8(g)} \quad \Delta H = x \text{ [kJ]}$$

❷ Write the chemical equations and enthalpy changes for the combustion reactions of the substances whose enthalpy values are given. Since the enthalpy of combustion of graphite is −394 kJ/mol,

$$C(\text{graphite}) + O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H = -394 \text{ kJ} \quad ......(i)$$

NOS: Philosophy of Science

Hess's Law demonstrates:

• Objectivity: Different scientists obtain the same ΔH when using the same data.

• Cumulativity: Hess law was built on earlier work concerning energy conservation to develop a more powerful method for calculating reaction enthalpies.

Since the enthalpy of combustion of hydrogen is −286 kJ/mol,

$H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(l)}$ $\Delta H = -286$ kJ ......(ii)

Since the enthalpy of combustion of propane is −2219 kJ/mol,

$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O(l)$ $\Delta H = -2219$ kJ ......(iii)

3 Combine the equations written in 2 to form the chemical equation in 1.

This can be derived using the relationship: (i)×3+(ii)×4−(iii).

4 Calculating the enthalpy of reaction in the same way as formed in 3,

$x = \{-394 \times 3 + (-286) \times 4\} - (-2219)$

$= -107$

Therefore, the required enthalpy of formation is −107 kJ/mol

(b) 1 Let x [kJ/mol] represent the bond enthalpy.

$H_2O_{(g)} \longrightarrow 2H_{(g)} + O_{(g)}$ $\Delta H = x$ [kJ]

2 Write the chemical equations and their corresponding enthalpy changes:

$H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(g)}$ $\Delta H = -242$ kJ ......(i)

$H_{2(g)} \longrightarrow 2H_{(g)}$ $\Delta H = 436$ kJ ......(ii)

$O_{2(g)} \longrightarrow 2O_{(g)}$ $\Delta H = 498$ kJ ......(iii)

3 The equation in 1 can be obtained by: (ii)+(iii)×$\frac{1}{2}$−(i).

4 Calculating using the same combination from 3: gives 436 kJ+249 kJ+242 kJ = 927 kJ

Since H_2O contains 2 O−H bonds, the bond enthalpy is: 927 kJ × $\frac{1}{2}$ = 463.5 kJ ≈ 464 kJ 464 kJ/mol.

Challenge your thinking

Many industrial reactions occur at temperatures and pressures that make direct enthalpy measurements difficult.

How can Hess's Law help in estimating energy changes before building a large-scale chemical plant?

....

Try

. Given the following enthalpies of formation:

- $CO_{2(g)}$: −394 kJ/mol
- $H_2O_{(g)}$: −286 kJ/mol
- $C_2H_5OH_{(l)}$: −277 kJ/mol.

(**a**) Write the chemical equation representing the combustion of ethanol, including the physical states.

Answer: ___________________________________

(**b**) Calculate the enthalpy of combustion of ethanol to the nearest integer.

Answer: ___________________________________

Exercise

Choose the correct answer:

1. Which scientific principle forms the basis of Hess›s law?

A. Conservation of mass

B. Conservation of energy

C. Le Chatelier›s principle

D. Boyle›s law

Answer: ___________________________________

2. Which statement best describes Hess›s law?

A. Enthalpy depends on reaction rate

B. Enthalpy depends on catalyst type

C. The total enthalpy change depends only on the initial and final states

D. Enthalpy depends on the reaction pathway.

Answer: ___________________________________

Answer on the following questions

3. Write the equation for the combustion of ethane $C_2H_{6(g)}$ including its enthalpy change, using the following data:

$C(graphite) + O_{2(g)} \longrightarrow CO_{2(g)} \quad \Delta H = -394\ kJ \quad \ldots(i)$

$H_{2(g)} + \frac{1}{2} O_{2(g)} \longrightarrow H_2O_{(l)} \quad \Delta H = -286\ kJ \quad \ldots(ii)$

$2C(graphite) + 3H_{2(g)} \longrightarrow C_2H_{6(g)} \quad \Delta H = -84\ kJ \quad \ldots(iii)$

Answer: ___________________________________

4. Calculate ΔH for the following reaction.

$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$

Given:

• Bond enthalpy of H−H: 436 kJ/mol

• Bond enthalpy of O=O: 498 kJ/mol

• Bond enthalpy of O−H: 463 kJ/mol

Challenge Your Thinking

$C(graphite) + O_{2(g)} \rightarrow CO_{2(g)} \quad \Delta H = -394\ kJ$

$H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)} \quad \Delta H = -286\ kJ$

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)} \quad \Delta H = -891\ kJ$

Determine the enthalpy of formation of methane:

$C(graphite) + 2H_{2(g)} \rightarrow CH_{4(g)}$

Explain every step using Hess's Law.

____

____

2-3 — Electrochemical Cell (1) (Daniell Cell)

✨ Engaging Introduction

Imagine that a power outage suddenly occurs in your city. Within seconds, millions of people depend on batteries, generators, and backup power systems to keep hospitals, communication networks, and transportation operating. But where does this electrical energy come from?

Surprisingly, electricity can be produced through simple chemical reactions involving metals and ions. Every battery, from a smartphone battery to an electric vehicle, relies on the same fundamental principle.

In this lesson, we will explore how chemical reactions can be transformed into electrical energy and how electrochemical cells power modern technology.

❓ **Guiding question:** How can a spontaneous chemical reaction be used to generate electrical energy that powers the devices and technologies we use every day?

Exploration

Materials

- Lemon, potato, or orange
- Copper coin or copper strip
- Galvanized nail (zinc-coated)
- Connecting wires
- Small LED

Procedure

1. Insert the zinc nail and copper strip into the fruit.
2. Connect wires to both metals.
3. Observe whether electrical energy is produced.
4. Test different fruits and compare performance.

⭐ **structure of science**

Key Fact :

An electrochemical cell converts chemical energy into electrical energy through a spontaneous redox reaction in which oxidation occurs at the anode and reduction occurs at the cathode.

Key Concepts:

· Electrochemical cell

· Anode

· Cathode

· Electromotive force (EMF)

· Daniell cell.

Inquiry Questions

• Which fruit produces the greatest voltage?

• Does changing the electrode metals affect performance?

• What evidence indicates that a redox reaction is occurring?

Point!

1. Electrochemical cell

(a) Electrochemical cell: A device that converts chemical energy into electrical energy using an oxidation-reduction (redox) reaction.

• The electrode from which electrons (e−) are released is called the anode (−), and the electrode that receives them is called the cathode (+).

(b) Reactivity series of metals: The arrangement of metals in order of their reactivity (or tendency to undergo oxidation).

$Li > K > Ca > Na > Mg > Al > Zn > Fe > Ni > Sn > Pb > (H_2) > Cu > Hg > Ag > Pt > Au$

(c) **Electromotive force (EMF)** : The voltage (potential difference) generated between the cathode and anode. The greater the difference in the reactivity series between the two metals used in the cell, the greater the EMF.

(d) Electrochemical Cell Notation

Represented in the form "(−) anode | Aqueous electrolyte solution | cathode (+)". This is called the cell notation .

Figure 2.3.1 metal samples

2. Daniell cell

Daniell cell : A galvanic cell consisting of a zinc plate immersed in an aqueous zinc sulfate ($ZnSO_4$) solution and a copper plate immersed in an aqueous copper(II) sulfate ($CuSO_4$) solution. The two half-cells are separated by a porous pot and connected by a conducting wire. (Figure on the right)

(a) Cell notation: (−) Zn | $ZnSO_{4(aq)}$ | $CuSO_{4(aq)}$ | Cu (+)

(b) Reactions at electrodes

• Reaction at the anode: $Zn \rightarrow Zn^{2+} + 2e^-$

• Reaction at the cathode: $Cu^{2+} + 2e^- \rightarrow Cu$

(c) Characteristics

• During discharge, ___Zn^{2+}___ from the aqueous zinc sulfate solution side and SO_4^{2-}___from the aqueous copper(II) sulfate solution side move through the porous pot to the opposite aqueous solutions.

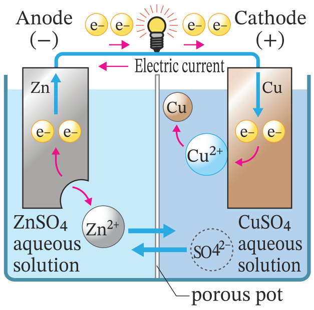


Figure 2.3.2 Daniell cell

Quick reference

Electrode	Sign	Process
Anode	−	oxidation — releases electrons ($Zn \to Zn^{2+} + 2e^-$)
Cathode	+	reduction — receives electrons ($Cu^{2+} + 2e^- \to Cu$)

Life Application

Every battery in a smartphone, laptop, smartwatch, or electric vehicle operates using electrochemical principles similar to those found in the Daniell cell.

Understanding electrochemical cells helps scientists develop:

• Longer-lasting batteries

• Faster charging systems

• More sustainable energy storage technologies

Be Global Citizen

Electrochemical cells play a major role in sustainable development. Where it uses in Electric vehicles, Renewable energy storage, Portable medical devices, Emergency power systems, Satellite technologies

Advances in battery technology can reduce fossil fuel dependence and support global sustainability goals.

Warm Up

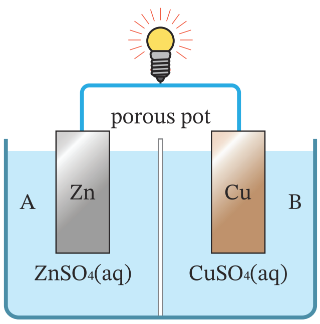
The figure is a schematic of a Daniell cell. Answer the following.

The figure on the right is a schematic diagram of a Daniell cell. Answer the following questions.

(**a**) Express the reactions occurring at the cathode (+) and anode (−) respectively,

using ionic equations containing electrons (e−).

(**b**) To maximise the discharge time, what should the initial concentrations of aqueous solutions A and B be? Answer using the terms ‘high concentration’ and ‘low concentration’.



Explain your reasoning by focusing on the ‘increase or decrease of ions’ at each electrode.

Explanation

(a) Cathode (+): $Cu^{2+} + 2e^- \to Cu$. Anode (−): $Zn \to Zn^{2+} + 2e^-$.

(b) [Concentration Condition]

The concentration of aqueous solution A ($ZnSO_4$) is set to be low, and the concentration of aqueous solution B ($CuSO_4$) is set to be high.

[Reason]

At the anode, Zn^{2+} ions are produced as the reaction proceeds. By starting with a low concentration of $ZnSO_4$ solution, it takes longer for the solution to reach saturation, allowing more zinc to dissolve. In contrast, at the cathode, Cu^{2+} ions are consumed. By using a high concentration of $CuSO_4$ solution initially, the depletion of Cu^{2+} ions is delayed, ensuring the reaction continues for a longer period.

⁉ Problem-Solving

Analyze Experimental Data

Two electrochemical cells are constructed:

Cell A: Zn | Cu

Cell B: Mg | Cu

Using the reactivity series:

- Predict which cell produces the higher EMF. ....
- Explain your reasoning. ....
- Identify the anode and cathode in each cell. ....

Try

Answer the following about the cell shown:

(a) What is the name of the cell shown in the figure on the right?

(b) Which metal plate is the cathode (+)? Answer with its chemical symbol.

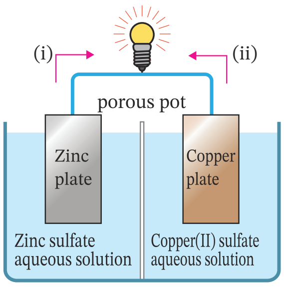
(c) When the zinc plate and the copper plate are connected by a conducting wire, express the changes that occur at both electrodes using ionic equations containing electrons (e−).

(d) Which direction, (i) or (ii) in the figure, do the electrons flow in the conducting wire?

(e) To maximize the total electrical charge delivered, how should the initial concentrations of the aqueous zinc sulfate solution and the aqueous copper(II) sulfate solution be chosen?

(f) In this cell, a potential difference is generated between the two electrodes. What is this potential difference called?

(g) If the zinc plate is replaced with a nickel plate and the aqueous solution is replaced with an aqueous nickel(II) sulfate solution, how will the potential difference change compared to that in (f)?



Answer: ____________________________________

Exercise

Choose the correct answer:

1. Which electrode gains mass during the operation of a Daniell cell?

A. Zinc electrode

B. Copper electrode

C. Both electrodes

D. Neither electrode

Answer: ___________________________________

2. In a Daniell cell, oxidation occurs at the:

A. Cathode

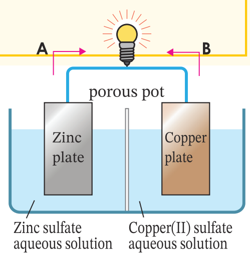
B. Electrolyte

C. Anode

D. Porous pot

Answer: ___________________________________

3. Answer the following questions about the cell shown in the figure on the right.



(a) Express the changes occurring at the anode (−) and cathode (+) respectively during discharge using reaction equations with electrons (e−).

(b) Which is the direction of the electric current, A or B in the figure?

(c) Which of the following moves from the aqueous copper(II) sulfate solution to the aqueous zinc sulfate solution through the porous pot? Answer with the corresponding number.

(i) Zn (ii) Zn^{2+} (iii) Cu (iv) Cu^{2+} (v) H^+ (vi) SO_4^{2-}

Challenge your thinking

A Daniell cell uses:

• Zn electrode in $ZnSO_4$ solution

• Cu electrode in $CuSO_4$ solution

A student proposes increasing the concentration of $ZnSO_4$ and decreasing the concentration of $CuSO_4$ to make the cell last longer.

Evaluate this proposal.

• Will it increase or decrease discharge time? ....

• Explain using changes in Zn^{2+} and Cu^{2+} concentrations at both electrodes.

....

2-4 — Electrochemical Cell (2) (Various Cells)

Engaging Introduction

A hydrogen-powered bus travels through a city without producing smoke, carbon dioxide, or harmful pollutants. Instead, the only substance released from its exhaust is water.At the same time, millions of cars around the world rely on lead-acid batteries to start their engines every day.

In this lesson, we will explore how different electrochemical cells store and produce electricity, and how scientists use chemical evidence to explain and predict their performance.

? Guiding question: How can chemical reactions be designed and controlled to provide electrical energy for different real-world applications?

Point!

1. Various cells

(a) Lead–acid battery: A battery using lead Pb for the anode, lead(IV) oxide PbO_2 for the cathode, and dilute sulfuric acid as the electrolyte.

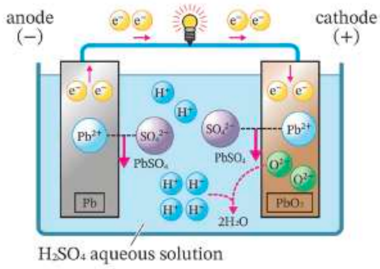


Figure 2.4.1 Lead–acid battery

(i) Cell notation: (−)Pb | H_2SO_4(aq) | PbO_2(+)

(ii) Reactions at electrodes (during discharge)

- Reaction at the anode: $Pb + SO_4^{2-} \longrightarrow PbSO_4 + 2e^-$
- Reaction at the cathode: $PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \longrightarrow PbSO_4 + 2H_2O$
- Overall reaction: $Pb + PbO_2 + 2H_2SO_4 \longrightarrow 2PbSO_4 + 2H_2O$

(iii) Characteristics

Because H2SO4 is used and H2O is produced

- The concentration of sulfuric acid becomes lower due to discharge.
- An external power source is connected to the lead-acid battery to charge it. During charging, reverse electrode reactions (opposite to discharge) occur at both the anode and cathode.

(b) Fuel cell : A device that extracts electrical energy using fuel (reducing agent) such as hydrogen, and an oxidant such as oxygen.

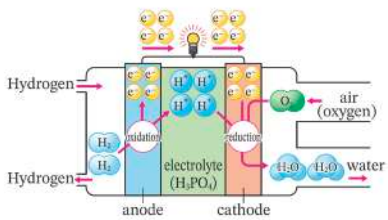


Figure 2.4.2 Fuel cell

(i) Cell notation (phosphoric acid type):

(−) H_2 | H_3PO_4(aq) | O_2 (+)

(ii) Reactions at electrodes (phosphoric acid type)

- Reaction at the anode: $H_2 \longrightarrow 2H^+ + 2e^-$
- Reaction at the cathode: $O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$
- Overall reaction: $2H_2 + O_2 \longrightarrow 2H_2O$

structure of science

Key Fact :

Electrochemical cells convert chemical energy into electrical energy through redox reactions, and the quantity of electricity produced depends on the amount of electrons transferred.

Key Concepts:

- Electrochemical cell
- Lead-acid battery
- Fuel cell
- Oxidation
- Reduction
- Faraday constant
- Electric charge
- Electrolyte
- Electrode reactions.

Think as an Engineer

Engineers designing electric vehicles must decide whether to use:

- Lead-acid batteries
- Lithium-ion batteries
- Hydrogen fuel cells

Challenge

Which technology would you choose for:

A school bus?..

A passenger car. ....

A backup power system? ....

Justify your choice using scientific evidence...

2. Quantity of electricity

(a) Expressing the quantity of electricity

The quantity of electricity when a current of 1 Ampere flows for (1 second) is expressed in a unit called 1 C (Coulomb).

(b) Faraday constant: The magnitude of the quantity of electricity per 1 mol of electrons. Denoted by the symbol F.

$$F = 9.65 \times 10^4 \text{ C/mol}$$

- The quantity of electricity that flow can be found using the Faraday constant.

Total electric charge [C] = Faraday constant [C/mol] × Amount of substance of electrons [mol]

(c) Chemical equations of cells

In a cell, since the reactions occurring at each electrode are different, consider the reaction equation to use according to what you want to find.

Key reactions & charge

Cell	Anode (−, oxidation)	Cathode (+, reduction)
Lead-acid	$Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$	$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$
Fuel cell (H_3PO_4)	$H_2 \rightarrow 2H^+ + 2e^-$	$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$

Life Application

Automobile Batteries

Lead-acid batteries provide the large burst of electrical energy needed to start vehicle engines.

Hydrogen-Powered Vehicles

Fuel cells generate electricity while producing only water as a by-product.

Warm Up

Answer the following. Assume F = 9.65×10^4 C/mol. (H = 1.0, O = 16, S = 32)

(a) Read the following text and answer the questions below. Assume the Faraday constant F = 9.65×10^4 C/mol.

(H = 1.0, O = 16, S = 32)

A lead-acid battery consists of (A) and lead immersed as electrode plates in dilute sulfuric acid. When 1.0 mol of lead completely changes during discharge, a quantity of electricity of (B) C can be extracted, and the dilute sulfuric acid decreases by (C) g.

Also, when charging the lead-acid battery, connect the lead to the (D) pole of the external power source. Upon charging, the mass of both electrode plates (E).

(i) Fill in the blanks ( ) with appropriate words and numerical values. Round to two significant figures.

(ii) Express the changes occurring during charging of the lead-acid battery together as a chemical equation.

(b) Read the following text and answer the questions below.

A fuel cell is a device that extracts the energy of a combustion reaction as electrical energy using fuel and oxygen. In a fuel cell with an aqueous phosphoric acid solution as the electrolyte, hydrogen is (A) at the (B) to become (C), which moves through the electrolyte and reacts with oxygen at the (D) to become (E).

(i) Fill in the blanks ( ) with the appropriate words.

(ii) Express the reactions at the anode and cathode using ionic equations with electrons e−.

⚠ Misconception Alert

Misconception

„Electricity is stored inside the wires."

Scientific Explanation

Electrical energy is generated by electron transfer during redox reactions.

The wires simply provide a pathway for electron flow.

Explanation

(a) (i) **A:** lead(IV) oxide

B: Since the amount of change in lead is known and the quantity of electricity is to be found, consider the reaction at the anode.

From the reaction at the anode, $Pb + SO_4^{2-} \longrightarrow PbSO_4 + 2e^-$, when 1.0 mol of lead changes, 2.0 mol of electrons flow.

Since Quantity of electricity flowed [C] = Faraday constant [C/mol]×Amount of substance of electrons [mol], 2.0 mol×9.65×10^4 C/mol = 1.93×10^5 C

$\approx 1.9\times10^5$ C

C: Since considering the amount of decrease in dilute sulfuric acid, consider the overall reaction.

$$Pb + PbO_2 + 2H_2SO_4 \longrightarrow 2PbSO_4 + 2H_2O$$

When 1.0 mol of lead changes, sulfuric acid decreases by 2.0 mol and water increases by 2.0 mol.

From H_2SO_4 = 98 and H_2O = 18, the amount of decrease in dilute sulfuric acid is,

(98−18) g/mol×2.0 mol = 160 = 1.6×10^2 g

D: Connect the anode and cathode of the lead-acid battery to the negative and positive terminals of the external power supply, respectively.

During charging, the negative terminal of the lead-acid battery (where lead is located) is connected to the negative terminal of the external power supply, acting as the cathode. negative.

E: During discharge, Pb and PbO_2 are converted to $PbSO_4$, increasing the electrode mass. During charging, the

reverse reactions occur, decreasing the electrode mass. decreases

(ii) During charging, the reaction in the reverse direction to discharge occurs.

Therefore, $2PbSO_4 + 2H_2O \longrightarrow Pb + PbO_2 + 2H_2SO_4$

(b) (i) **A**: oxidised **B**: anode **C**: hydrogen ions **D**: cathode E: water

(ii) Anode: $H_2 \longrightarrow 2H^+ + 2e^-$

Cathode: $O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$

Identifying variables

A scientist wants to investigate how the amount of electrical charge produced by a lead-acid battery depends on the amount of lead that reacts during discharge. Several identical lead-acid batteries are prepared. The scientist changes the amount of lead available in each battery while keeping all other conditions the same. The total electrical charge produced is then measured.

1. What is the independent variable? ....

2. What is the dependent variable? ....

3. Which variables should be controlled to ensure a fair test? ....

Try

Choose the correct answer:

1. What happens to the sulfuric-acid concentration during discharge of a lead-acid battery?

A. Increases B. Remains constant

C. Decreases D. Doubles

Answer: ____________________________________

2. Which substance is produced at the cathode of a hydrogen fuel cell?

A. Hydrogen gas B. Sulfuric acid

C. Water D. Oxygen gas

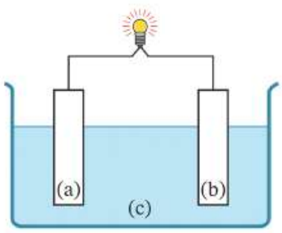
Answer: ____________________________________

3. Answer the following questions regarding the lead-acid battery.

Assume the Faraday constant $F = 9.65 \times 10^4$ C/mol.
(H = 1.0, O = 16, S = 32, Pb = 207)

(a) What substances are used for the (a) anode (−), (b) cathode (+), and (c) electrolyte? Answer each with its chemical formula.

(b) Express the reactions at the anode (−) and cathode (+) during discharge using ionic equations including e−,



respectively.

(c) Express the overall reaction of the battery during discharge as a chemical equation.

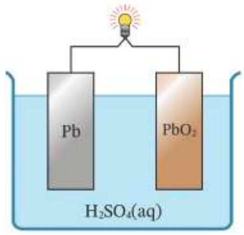
(d) How does the concentration of the electrolyte change due to discharge?

(e) After discharging this battery for a certain period, it was charged, and the mass of the anode decreased by 7.2 g compared to before charging. What is the quantity of electricity in C that flowed during charging? Give your answer to two significant figures.

Exercise

1. The figure below is a schematic diagram of a lead-acid battery. Answer the following questions.

(H = 1.0, O = 16, S = 32, Pb = 207)



(a) Name the substance used for the cathode (+).

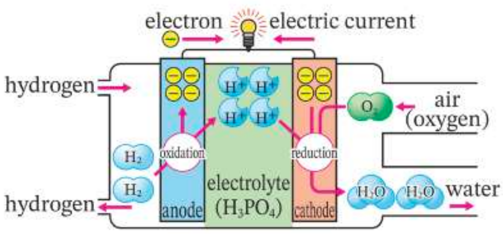
(b) Express the reactions at the cathode (+) and anode (−) during discharge using ionic equations including electrons e−, respectively.

(c) Express the overall reaction of the battery during discharge as a chemical equation.

(d) How does the concentration of the electrolyte change due to discharge?

(e) When a quantity of electricity equivalent to 1.0 mol of electrons flows, by how many grams does the mass of the positive electrode increase or decrease? Give your answer to two significant figures.

2. The figure on the right is a schematic diagram of a fuel cell (phosphoric acid type). A fuel cell extracts the energy generated when hydrogen (fuel) and oxygen react as electrical energy. Answer the following questions.



(a) Regarding the fuel cell (phosphoric acid type), answer the (a) reaction at the cathode (+), (b) reaction at the anode (−), and (c) overall reaction. For (a) and (b), answer with ionic equations including electrons e−, and for (c), answer with a reaction equation.

(b) Calculate the quantity of electricity that flows when 5.6 L of hydrogen is consumed at 0 °C and 1.013×10^5 Pa. Assume sufficient oxygen is present. (Faraday constant: $F = 9.65 \times 10^4$ C/mol) Express your answer to two significant figures.

Challenge Your Thinking

A lead-acid battery discharges completely and causes 3.0 mol of electrons to flow.

Using:

$F = 9.65 \times 10^4$ C/ mol

Calculate:

1. Total charge produced...

2. Whether the charge would be greater or smaller if 5.0 mol of electrons flowed.

....

3. Explain your reasoning. ....

2-5 — Electrolysis (1) (Reactions at Electrodes)

Engaging Introduction

A company wants to produce hydrogen fuel for clean energy and chlorine for water treatment. Instead of adding chemicals, engineers simply pass an electric current through a salt solution. A few minutes later, bubbles appear at both electrodes, and entirely new substances are formed.

Let`s explore in this lesson, How can electricity transform substances into new products?.

? Guiding question: How can electrical energy force chemical reactions to occur, and how do chemists predict which substances will be produced at each electrode?

Point!

1. Electrolysis

(a) Electrolysis: A process that uses electrical energy to drive non-spontaneous redox reactions.

• The reaction at the electrode changes depending on the type of electrode and the type of ions in the aqueous solution.

(b) Reactions at electrodes

《**Reactions at the anode**》

(i) When the electrode is platinum or carbon

• When halide ions (such as Cl^- and I^-) are present

Halide ions are oxidised to form elements . Example: $2I^- \longrightarrow I_2 + 2e^-$

• When other ions (such as sulfate ions SO_4^{2-} and nitrate ions NO_3^-) are included

... When the aqueous solution is acidic or neutral, water H_2O is oxidised to generate O_2.

$2H_2O \longrightarrow O_2 + 4H^+ + 4e^-$

... When the aqueous solution is basic, hydroxide ions OH^- are oxidised to generate O_2.

$4OH^- \longrightarrow O_2 + 2H_2O + 4e^-$

(ii) When the electrode is other than platinum or carbon

The metal of the electrode is oxidised and dissolves as cations.

Example: $Cu \longrightarrow Cu^{2+} + 2e^-$

《Reactions at the cathode》 The type of electrode does not matter.

(i) When metal ions with low reactivity (such as Cu^{2+} and Ag^+) are present at the cathode, they are reduced and the corresponding metals are deposited.

Example: $Ag^+ + e^- \longrightarrow Ag$

(ii) When metal ions with a higher reactivity than hydrogen are present, water is preferentially reduced at the cathode rather than the metal ions, generating hydrogen gas.

$2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$

structure of science

Key Fact :

The products of electrolysis depend on both the ions present in solution and the type of electrode used.

Key Concepts:

· Electrolysis

· Reactions at electrodes

· Cation exchange membrane

· Ion exchange membrane method.

(iii) When the aqueous solution is acidic, hydrogen ions (H+) are reduced at the cathode to produce hydrogen gas.

$2H^+ + 2e^- \longrightarrow H_2$

The electrodes are platinum, sulfate ions (SO_4^{2-}) are present, and the aqueous solution is acidic.

(c) Example of electrolysis

⟨Electrolysis of aqueous sulfuric acid (H_2SO_4) solution (dilute sulfuric acid (Electrode: Platinum)⟩ ●

(i) Anode: $2H_2O \longrightarrow O_2 + 4H^+ + 4e^-$

(ii) Cathode: $2H^+ + 2e^- \longrightarrow H_2$

Inquiry Exploration

Cell	Electrolyte	Electrode
A	NaCl(aq)	Carbon
B	$AgNO_3$(aq)	Silver
C	NaOH(aq)	Platinum

1. Which ions are present in each solution? ....

2. Which ions are most likely to gain electrons? ....

3. Which particles are most likely to lose electrons? ....

2. Ion exchange membrane method

(a) Cation exchange membrane: A membrane through which only cations can pass.

(b) Ion exchange membrane method: A manufacturing method for sodium hydroxide using a cation exchange membrane.

❶ Electrolyze an aqueous sodium chloride solution using carbon and iron electrodes.

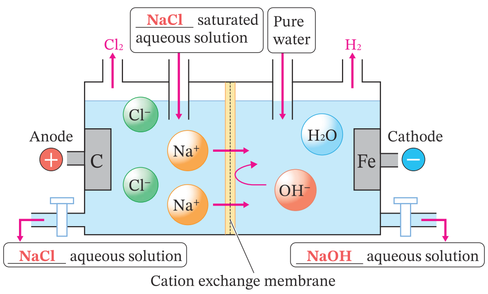
- Reaction at the anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$
- Reaction at the cathode: $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$

Cl_2 is generated on the anode side, and H_2 is generated on the cathode side.

❷ Passing through the cation exchange membrane, Na+ moves to the cathode side.

❸ On the cathode side, the concentrations of OH− and Na+ increase.

→ High-purity sodium hydroxide is obtained by concentrating the solution from the cathode compartment.



Life Application

The Hydrogen Economy: Society currently depends on fossil fuels (coal, oil, gas), but burning them releases CO_2, driving the greenhouse effect and rising sea levels. Hydrogen (H_2) is an ideal alternative because its combustion produces only water, making it a non-polluting, renewable fuel. The "hydrogen economy" rests on two electrochemical technologies: (i) electrolysis of water — using solar energy to split water into hydrogen; (ii) fuel cells — converting the chemical energy of hydrogen back into electricity. Mastering electrochemistry is a key to a sustainable, carbon-free future.

Warm Up

The solutions (a)–(c) were electrolysed. Write the half-reactions at each electrode.

	(a)	(b)	(c)
Electrolyte	copper(II) chloride solution	silver nitrate solution	sodium hydroxide solution
Anode	C	Ag	Pt
Cathode	C	Ag	Pt

Explanation

(a) Anode: The electrode is carbon, and copper(II) chloride ionizes, providing chloride ions (Cl^-) in the solution. These chloride ions are then oxidised to form elemental chlorine. Therefore, $2Cl^- \longrightarrow Cl_2 + 2e^-$

Cathode: Copper(II) chloride ionizes, and ions of less reactive metals, Cu^{2+}, are present. At this time, the metal ions are reduced and metal is deposited.

Therefore, $Cu^{2+} + 2e^- \longrightarrow Cu$

(b) Anode: Since the electrode is silver, the metal of the electrode is oxidised and dissolves as cations.

Therefore, $Ag \longrightarrow Ag^+ + e^-$

Cathode: Silver nitrate ionizes, and less reactive metal ions, Ag+, are present.

Therefore, $Ag^+ + e^- \longrightarrow Ag$

(c) Anode: The electrode is platinum, and the aqueous sodium hydroxide (NaOH) solution is basic. At this time, hydroxide ions OH− are oxidised.

Therefore, $4OH^- \longrightarrow O_2 + 2H_2O + 4e^-$

Cathode: Sodium hydroxide ionizes, and ions of metals more reactive than hydrogen, Na+, are present. At this time,

water H_2O is reduced to generate H_2.

Therefore, $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$.

Think Like a Chemist

A student says: "Sodium metal should form during the electrolysis of sodium chloride solution because sodium ions are present".

Evaluate this statement.

• Is sodium more reactive than hydrogen?

....

• Which species is easier to reduce?

....

• What evidence supports your conclusion?

....

Try

Solve the Following Questions

1. The aqueous electrolyte solutions in the table were electrolyzed. Express the changes A to F occurring at each electrode using half-reaction equations.

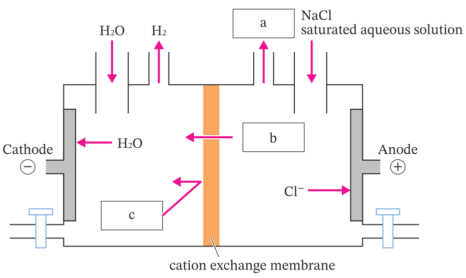
Aqueous electrolyte	Cathode	Change	Anode	Change
(a) Dilute sulfuric acid	Cu	(A)	Cu	(B)
(b) Sodium hydroxide solution	Pt	(C)	Pt	(D)
(c) Copper(II) sulfate solution	Pt	(E)	Pt	(F)

Answer: ___________________________________

2. The figure below schematically shows the industrial production of sodium hydroxide by the ion exchange membrane method.

Choose the most appropriate combination of chemical formulas to insert into a to c in the figure from (i) to (vi) below.

	a	b	c
(i)	O_2	Na^+	OH^-
(ii)	O_2	OH^-	OH^-
(iii)	O_2	Na^+	Cl^-
(iv)	Cl_2	OH^-	Cl^-
(v)	Cl_2	Na^+	Cl^-
(vi)	Cl_2	Na^+	OH^-



Answer: ___________________________________

Exercise

Choose the correct answer:

1. During electrolysis of $AgNO_3(aq)$, silver metal deposits at the cathode because:

A. Ag^+ ions are easily reduced

B. Water is oxidized

C. Nitrate ions gain electrons

D. Silver is highly reactive

Answer: ___________________________________

2. Which factor would most likely change the products formed during electrolysis?

A. Colour of the container

B. Electrode material

C. Shape of the beaker

D. Position of the power supply

Answer: ___________________________________

Answer on the Following Questions

3. When electrolysis is performed using the combinations of electrolyte and electrode from (a) to (d) in the table, express the reactions occurring at the anode and cathode using half-reaction equations, respectively.

	(a)	(b)	(c)	(d)
Electrolyte	Dilute sulfuric acid	KOH solution	$CuSO_4$ solution	NaCl solution
Anode	Pt	Pt	Cu	C (graphite)
Cathode	Pt	Pt	Fe	Fe

Answer: ___________________________________

Challenge Your Thinking

A concentrated aqueous solution of sodium chloride is electrolyzed using carbon electrodes.

1. Identify all ions present. ....

2. Predict the product at the anode. ....

3. Predict the product at the cathode. ....

4. Explain why sodium metal is not produced. ....

5. Predict how the pH of the cathode compartment changes ....

2-6 — Electrolysis (2) (Quantity of Electricity, Two Electrolytic Cells)

✨ Engaging Introduction

A company uses electrolysis to produce pure copper for electrical wires. Engineers know exactly how much copper will be deposited before the process even begins—not by weighing the copper, but by measuring the electric current flowing through the circuit.

In this lesson, we will explore the relationship between electricity, electrons, and chemical change, and discover how chemists use quantitative evidence to predict industrial production.

> ❓ **Guiding question:** How can chemists predict the amount of substance produced during electrolysis using only electrical measurements such as current and time?

Exploration

Students can electroplate a coin or a metal key using copper sulfate solution.

Observation

Before electrolysis:

....

After electrolysis:

....

Inference

....

Point!

1. Electrolysis and quantity of electricity

(a) Calculating the Quantity of Electricity Quantity of electricity (C) = Current (A)×Time (s)

(b) Faraday constant F: The magnitude of the quantity of electricity per 1 mol of electrons. 9.65×10^4 C/mol.

(c) the mass of a substance (m) deposited or liberated at an electrode is directly proportional to the quantity of electricity (Q) passed through the electrolyte.

- Expression: m α Q
- Since Q=I×t (Current × Time), we can write: m=ZIt
- m = mass of substance deposited (in grams)
- Q = quantity of electricity passed (in Coulombs)
- I = current (in Amperes)
- t = time (in seconds)
- Z = Electrochemical Equivalent of the substance.

⭐ Structure of Science

Key Fact :

The amount of substance produced during electrolysis is directly proportional to the quantity of electricity that passes through the circuit.

Key Concepts:

· Electrolysis
· Quantity of electricity
· Faraday constant
· Oxidation
· Reduction
· Electrolytic refining.

The Electrochemical Equivalent (Z) is the mass of a substance deposited by one Coulomb of charge. Its value is given by: Z=molar mass/nF

- n = number of electrons transferred per mole of the substance.
- F = Faraday constant ($9.65x10^4$ C), which is the charge of one mole of electrons.

2. Electrolysis using two electrolytic cells

Even if electrolytic cells are connected, consider electrolysis for each electrolytic cell one by one. Electrons move through the two electrolytic cells.

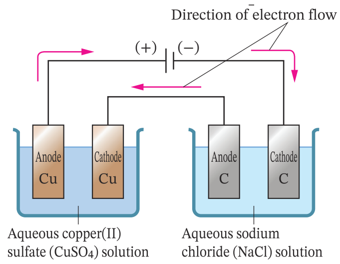


Figure 2.6.1 Electrolysis using two electrolytic cells

(a) Reactions at electrodes (in the case of the figure on the right. 2.6.1)

(i)Reaction in aqueous $CuSO_4$solution

The electrodes are copper Cu, Cu^{2+} is present in the aqueous solution.

Anode: $Cu \longrightarrow Cu^{2+} + 2e^-$

Cathode: $Cu^{2+} + 2e^- \longrightarrow Cu$

(ii)Reaction in aqueous NaCl solution

The electrodes are carbon, Cl^- and Na^+ are present in the aqueous solution.

Anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$

Cathode: $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$

(b) Circuit and magnitude of quantity of electricity

When electrolytic cells A and B are connected in series, letting the quantity of electricity flowing through electrolytic cells A and B be Q_A and Q_B, respectively, and the quantity of electricity flowing through the entire circuit be Q, the relationship Q = $Q_A = Q_B$ holds.

Quantity of electricity [C] = Current [A] × Time [s]. In a series circuit, the current flowing through the circuit is the same everywhere.

3. Electrolytic refining of Copper

(a) Electrolytic refining: Obtaining a pure metal from a metal containing impurities using electrolysis.

(b)Electrolytic refining of copper

By smelting chalcopyrite, which is a copper ore, to obtain “impure copper” with a purity of about 99%, and further electrolytically refining it, “pure copper” with a purity of 99.99% or more is obtained.

⟨Mechanism of electrolytic refining of copper⟩ (Figure on the right 2.6.2)

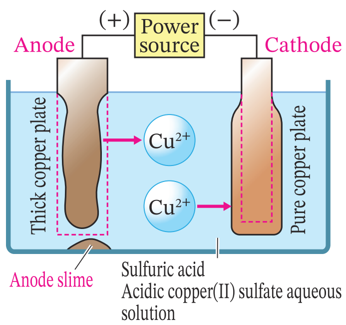


Figure 2.6.2. Electrolytic Refining of Copper

❶ Make the impure copper plate the anode(+) and the pure copper plate the cathode(−) .

❷ At the anode, copper atoms lose electrons and dissolve as copper(II) ions.

Chemical equation: $Cu \longrightarrow Cu^{2+} + 2e^-$

At the cathode, copper(II) ions receive electrons and precipitate as copper.

Chemical equation: $Cu^{2+} + 2e^- \longrightarrow Cu$

❸ At the anode, gold, silver, and other less reactive metals contained in the impure copper precipitate as elements. This precipitate is called anode slime .

Quantity of electricity — key relations

Relation	Formula
Charge from current & time	Q [C] = I [A] × t [s]
Charge from electrons	Q [C] = F × n(e^-), F = 9.65 × 10^4 C/mol
Cells in series	Q = Q_A = Q_B (same electrons through each cell)
Copper refining	impure Cu = anode (+), pure Cu = cathode (−); Au/Ag → anode slime

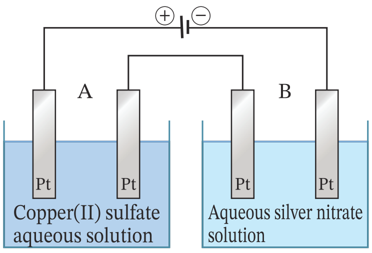
Life Application

Electrolytic refining ensures Better electrical conductivity, Reduced energy loss, More reliable electronic devices Without electrolysis, modern communication systems would not function efficiently.

Warm Up

Answer the following questions

(a) As shown in the figure on the right, an aqueous copper(II) sulfate solution was placed in electrolytic cell A, an aqueous silver nitrate solution in electrolytic cell B, and electrolysis was performed using platinum as electrodes. When a constant current was passed, the mass of the cathode of electrolytic cell B increased by 5.4 g. Assuming the Faraday constant $F = 9.65 \times 10^4$ C/mol, answer the following questions. Provide answers to two significant figures. (Cu = 64, Ag = 108)



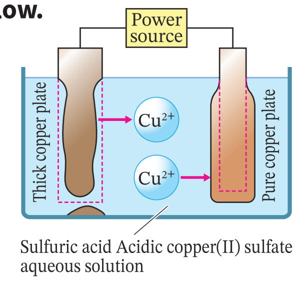
(i)Write the reaction at the cathode of electrolytic cell B as a half-reaction equation.

(ii)What quantity of electricity in C is required for this electrolysis?

(iii)What mass of copper in g was deposited on the cathode of electrolytic cell A?

(b) **Read the following text and answer the questions below.**

Copper extracted from chalcopyrite is called impure copper and contains impurities such as gold, silver, iron, and zinc. By making impure copper the (**A**) electrode and pure copper the (**B**) electrode, electrolysis is performed in an aqueous copper(II) sulfate solution to obtain pure copper. At this time, metals with a (**C**) reactivity such as gold and silver precipitate as elements at the bottom of the electrode, and this precipitate is called (**D**).



(i)Answer the words that apply to A to D.

(ii)Express the reactions at the anode and cathode using half-reaction equations.

Explanation

(a) (i) In the aqueous solution of electrolytic cell B, metal ions of a less reactive metal,

Ag^+, are present. $Ag^+ + e^- \longrightarrow Ag$

(ii) Consider the reaction at the cathode of electrolytic cell B, for which the mass increase is known.

From the reaction equation in (i), it can be seen that when 1 mol of Ag is formed, 1 mol of electrons flows. From the mass increase of the cathode of electrolytic cell B, 5.4 g of Ag is formed, so the amount of substance of Ag formed is

$\frac{5.4\ g}{108\ g/mol} = 5.0\times10^{-2}$ mol

Therefore, the moles of electrons that flowed is also 5.0×10^{-2} mol. Thus, the required quantity of electricity is,

9.65×10^4 C/mol$\times5.0\times10^{-2}$ mol $= 4.825\times10^3 \approx 4.8\times10^3$C

(iii)The reaction at the cathode of electrolytic cell A is $Cu^{2+} + 2e^- \longrightarrow Cu$.

From the reaction equation, it can be seen that when 1 mol of Cu is formed, 2 mol of electrons flow.

In a series circuit, since the current flowing through the circuit is the same everywhere, the moles of electrons flowing is also the same.

The electrons that flowed into electrolytic cell A are 5.0×10^{-2} mol, and the amount of substance of Cu formed is $5.0\times10^{-2}\times\frac{1}{2}$ mol.

Therefore, the mass of the deposited copper is,

64 g/mol$\times5.0\times10^{-2}\times\frac{1}{2}$ mol $= 1.6$ g

(b) (i) **A**: anode **B**: cathode **C**: low **D**: anode slime

(ii) Anode: $Cu \longrightarrow Cu^{2+} + 2e^-$

Cathode: $Cu^{2+} + 2e^- \longrightarrow Cu$

NOS: Identifying variables

A scientist wants to investigate the effect of electric current on the mass of copper deposited during electrolysis.

1. What is the independent variable?

....

2. What is the dependent variable?

....

3. Which variables should be controlled to ensure a fair test?

....

Try

Choose the correct answer:

1. Which equation directly relates current and time to the quantity of electricity?

A. PV = nRT B. Q = It

C. π = cRT D. Δt = Km

Answer: ___________________________________

2. In copper electrorefining, impure copper acts as the:

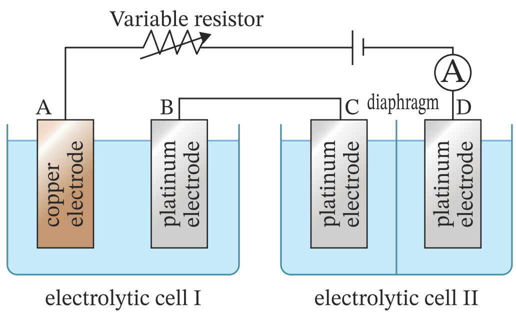
A. Cathode B. Electrolyte

C. Anode D. Salt bridge

Answer: ___________________________________

Answer on the following question

3. An electrolytic apparatus was assembled as shown in the figure, and an aqueous copper(II) sulfate solution was placed in electrolytic cell I, and an aqueous sodium sulfate solution in electrolytic cell II. Using this apparatus, electrolysis was performed for 32 minutes and 10 seconds while maintaining the current at 10 A. Answer the following questions. Assume the Faraday constant F = 9.65×10^4 C/mol.



(a) How many moles of electrons flowed?

(b) Express the changes occurring at electrodes A and B using half-reaction equations.

(c) Which is the closest value to the mass change [g] of electrode A in electrolytic cell I after electrolysis? (Cu = 64)

A −19.2 B −12.8 C −6.4 D ±0 E +6.4 F +12.8 G +19.2

(d) What is the volume in L of the gas generated at electrode C of electrolytic cell II at 0°C and 1.013×10^5 Pa? Answer to two significant figures.

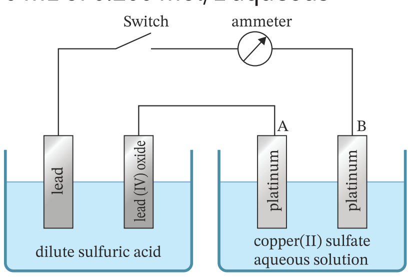
(e) What property does the aqueous solution near electrode D in electrolytic cell II exhibit after electrolysis?

Exercise

Read the following text and answer the questions below. Assume the Faraday constant **F** = *9.65×10^4 C/mol.*

(H = 1.0, O = 16, S = 32, Cu = 64, Pb = 207)

As shown in the figure, a lead-acid battery, an electrolytic cell, and an ammeter were connected in series. The electrolytic cell contains 200 mL of 0.200 mol/L aqueous copper(II) sulfate solution. Platinum plates were used for the electrodes of the electrolytic cell, and electrolysis was performed with a current of 2.00 A for 32 minutes and 10 seconds. In the electrolytic cell, the platinum plate A in the figure acts as the anode, and the platinum plate B acts as the cathode. On electrode B, (i) was deposited.



The lead-acid battery used for electrolysis is a battery in which dilute sulfuric acid with a density of 1.20 g/cm^3 and a concentration of 30.0% by mass is used, and lead is immersed as the anode (−) and lead(IV) oxide as the cathode (+).

(I)At the lead electrode, a(ii) reaction occurs, and (II)at the lead(IV) oxide electrode, a (iii) reaction occurs. When discharge continues, white (iv) is formed on both electrodes, and the electromotive force decreases. It can be charged by connecting the electrodes of a discharged lead-acid battery to the electrodes of a separate power source and passing electricity.

(a) Choose the appropriate words to enter in ( ) in the text, one by one, from the following words. The same word may be used repeatedly.

Oxidation, Reduction, Copper, Lead, Copper(II) sulfate, Lead(II) sulfate, Copper(II) oxide, Lead(IV) oxide

(b) Write the half-reactions for the underlined parts (I) and (II), showing electron transfer.

(c)Using platinum electrodes, write the half-reaction at electrode A (part 1). Then, when copper electrodes replace the platinum electrodes, write the half-reaction at electrode A (part 2)

(d) Calculate the change in mass of the lead anode (−) of the lead-acid battery, including increase or decrease, to the first decimal place.

Challenge Your Thinking

Two electrolytic cells are connected in series.

• Cell A contains $CuSO_4$ solution.

• Cell B contains $AgNO_3$ solution.

During electrolysis, 10.8 g of silver is deposited in Cell B.

Without directly using the example from the lesson:

1. Determine the moles of silver produced. ……………………………………………………

2. Calculate the moles of electrons transferred. ………………………………………………

3. Calculate the mass of copper deposited in Cell A.…………………………………………

Lesson 2-7 — Rate of Reaction

Engaging Introduction

A piece of iron left outdoors may take months or years to rust, while a tablet dropped into water can begin reacting almost immediately. In industry, some reactions must occur rapidly to increase production, whereas others must be slowed down for safety reasons.

In this lesson, we will explore how reaction rates are measured and how mathematical models help scientists understand the factors that control chemical change.

? Guiding question: Why do some chemical reactions occur in seconds while others take days, months, or even years, and how can chemists predict the speed of a reaction?

Point!

1. The Concept of Reaction Rate

Chemical reactions vary widely, ranging from **fast reactions** that complete instantly to **slow reactions** that proceed over a very long time.

• **Example of a fast reaction:** The formation of a white precipitate of silver chloride when an aqueous solution of sodium chloride is added to an aqueous solution of silver nitrate.

• **Example of a slow reaction:** The slow oxidation of iron by atmospheric moisture and oxygen outdoors, leading to the formation of rust.

Even the same reaction can change its rate depending on conditions such as temperature and pressure.

NOS: Identifying variables

Students investigate how temperature affects reaction rate by observing the reaction between an effervescent tablet and water.

1. What is the independent variable?

....

2. What is the dependent variable?

....

3. Which variables should be controlled to ensure a fair test?

....

2. Definition and Expression of Reaction Rate

Rate of Reaction: The decrease in the amount of reactant or the increase in the amount of product per unit time. The change in the amount of substance is usually considered as the change in concentration.

$$\text{Rate of reaction } v = \frac{\text{Amount of decrease in molar concentration of reactants}}{\text{Reaction time}} \text{ or }$$

$$\frac{\text{Amount of increase in molar concentration of products}}{\text{Reaction time}}$$

structure of science

Key Fact :

Reaction rate depends on reactant concentration and temperature, and its relationship is described by an experimentally determined rate equation.

Key Concepts:

· Reaction rate
· Concentration
· Rate equation
· Rate constant
· Reaction order.

3. Ratio of Reaction Rates and Stoichiometric Coefficients

Although the change in the molar concentration per unit time differs depending on **the substance focused on in a chemical reaction, the ratio of the rates of reaction for each substance is** equal to the ratio of their stoichiometric coefficients in the balanced chemical equation.

Example: When hydrogen and iodine are placed in a closed container and heated to a constant temperature, the following reaction occurs, generating hydrogen iodide:

$H_2 + I_2 \rightarrow 2HI$

The rate of disappearance of H_2 or I_2 : the rate of formation of HI = 1 : 2.

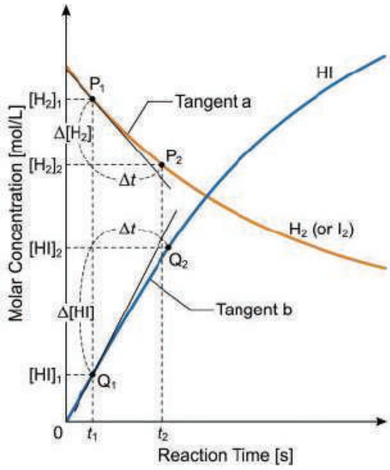


Figure 2.7.1 Reaction Rate and Stoichiometric Coefficients

4. The Rate Equation

The rate equation expresses the relationship between the rate of reaction and the molar concentrations of the substances involved in the reaction.

For a reaction represented by the chemical equation A + B → C + D, assuming the rate of reaction is v, the molar concentrations of reactants A and B are [A] and [B], and k, a, and b are constants, the rate equation (rate law) can be expressed as follows:

$v = k\,[A]^a\,[B]^b$

• **The Rate Constant** (k): This is the proportionality constant in the rate equation. The value of k is constant if the temperature is constant, but it changes if the temperature changes.

• **The Order of Reaction** (a, b): These are the orders of reaction with respect to reactants A and B, respectively. These values **must be determined experimentally** and do not necessarily match the stoichiometric coefficients in the chemical equation.

Reference table

Quantity	Expression	Notes
Rate of reaction	$v = \frac{-\Delta[\text{reactant}]}{\Delta t} = \frac{+\Delta[\text{product}]}{\Delta t}$	change per unit time; usually as a change in concentration
Rate ratio	rate ∝ stoichiometric coefficient	e.g. for $H_2 + I_2 \rightarrow 2HI$ the rates $H_2 : I_2 : HI$ = 1 : 1 : 2
Rate equation (rate law)	$v = k\,[A]^a\,[B]^b$	a, b are reaction orders — found experimentally
Rate constant k	proportionality constant	constant at fixed T; changes when T changes

Life Application

Refrigerators and Freezers: Placing perishable food in the fridge drops the temperature, which significantly slows down the chemical and biological reactions that cause food to spoil or bacteria to multiply.

Warm Up

Answer the following questions.

(a) Consider the reaction represented by A + B → C. By varying the concentrations of A and B, the instantaneous rate of reaction v was found for each, and the results shown in the table below were obtained.

Answer the questions below.

Experiment	[A] [mol/L]	[B] [mol/L]	v [mol/(L·s)]
1	0.30	1.20	3.6×10^{-2}
2	0.30	0.60	9.0×10^{-3}
3	0.60	0.60	1.8×10^{-2}

(i) Express the relationship between [A], [B], and v as a rate equation, letting the rate constant be k.

(ii) Give the value of the rate constant k to two significant figures.

(b) For a certain reaction, the rate constant doubles when the temperature rises by 10 K near room temperature. In this reaction, when the reaction temperature is changed from 25°C to 55°C, by what factor does the rate v increase?

Explanation

(a)(i) From experiment 1 and experiment 2, when [B] is halved, v becomes $\frac{1}{4}$, so v is proportional to the square of [B].

Also, from experiment 2 and experiment 3, when [A] is doubled, v is doubled, so it can be seen that v is proportional to the first power of [A]. Therefore, $\underline{v = k\,[A]\,[B]^2}$

(a)(ii) Once the rate equation is determined, the rate constant k can be found from the experimental results.

From the results of experiment 3,

$1.8\times10^{-2}\ \mathrm{mol/(L\cdot s)} = k \times (0.60\ \mathrm{mol/L}) \times (0.60\ \mathrm{mol/L})^2$

$k = \dfrac{1.8\times10^{-2}\ \mathrm{mol/(L\cdot s)}}{0.60\ \mathrm{mol/L} \times (0.60\ \mathrm{mol/L})^2}$

$= 0.0833\ \mathrm{L^2/(mol^2\cdot s)}$

$\approx 8.3\times10^{-2}\ \mathrm{L^2/(mol^2\cdot s)}$

Therefore, the value of k is $\underline{8.3\times10^{-2}}$

(b) Changing from 25°C to 55°C increases the temperature by 30 K.

Since the rate constant doubles for every 10 K rise in temperature, when the temperature rises by 30 K, the rate constant increases by a factor of $2^3 = 8$.

Therefore, <u>the rate constant becomes 8 times the original value</u>.

Think Like a Chemist

A pharmaceutical company develops a medicine that decomposes too quickly during storage.

Questions

- Should the medicine be stored at higher or lower temperatures? ....
- Which factors could slow the decomposition reaction? ....
- How could reaction-rate knowledge improve product safety? ....

Try

1. When 10.0 mL of 1.00 mol/L hydrogen peroxide solution is added to a small amount of manganese(IV) oxide, the hydrogen peroxide decomposes and oxygen is generated. The concentration of hydrogen peroxide became 0.64 mol/L 2 minutes after the start of the reaction. Answer the following questions to two significant figures.

(a) What is the rate of decomposition of hydrogen peroxide in mol/(L·s) during this time?

(b) What is the rate of generation of oxygen in mol/s?

Answer: ____________________________________

2. For a reaction that forms product C from reactants A and B at 25 °C, the rate of formation of C, v, quadrupled when only (A) was doubled, and decreased to ¼ of its original value when only (B) was halved (½). Answer the following questions.

(a) Letting the rate constant be k, write the rate equation expressing the relationship between v, (A), and (B).

Answer: ____________________________________

(b) When both (A) and (B) are doubled, by what factor does v increase?

Answer: ____________________________________

3. Consider a reaction where C is formed from A and B. When the rate of formation v [mol/(L·s)] of C was found by changing the molar concentrations [A] and [B], the results in the table below were obtained.

Experiment	[A] [mol/L]	[B] [mol/L]	v [mol/(L·s)]
(i)	0.30	0.60	0.0090
(ii)	0.30	1.20	0.036
(iii)	0.60	0.60	0.018

(a) Write the rate equation expressing the relationship between [A], [B], and v. Assume the proportionality constant is k.

Answer: ____________________________________

(b) What is the proportionality constant k called?

Answer: ____________________________________

(c) Find the value of the proportionality constant k to two significant figures.

Answer: ____________________________________

Exercise

Choose the correct answer:

1. For the reaction $H_2 + I_2 \rightarrow 2HI$, the rate of formation of HI compared with the rate of disappearance of H_2 is:

A. 1 : 1

B. 2 : 1

C. 1 : 2

D. 4 : 1

Answer: ___________________________________

2. Which quantity changes when temperature changes?

A. Reaction order

B. Stoichiometric coefficient

C. Rate constant (k)

D. Balanced equation

Answer: ___________________________________

Answer on the following question

3. For a reaction that forms product C from reactants A and B at 25 °C, the rate of formation of C, v, doubled when only [A] was doubled, and decreased to ¼ of its original value when only [B] was halved (½). Answer the following questions.

(a) Write the unit of v. Assume the unit of time is minutes [min].

Answer: ___________________________________

(b) Letting the rate constant be k, write the rate equation expressing the relationship between v, [A], and [B].

Answer: ___________________________________

(c) When [A] is tripled and [B] is quadrupled, by what factor does v increase?

Answer: ___________________________________

Challenge Your Thinking

A reaction follows the rate law: $v = k[A][B]^2$ A student doubles

both reactants simultaneously.

Challenge Questions - By what factor does the reaction rate increase?

Answer: ___________________________________

- Which reactant has the greater effect on reaction rate?

Answer: ___________________________________

- Predict what would happen if [A] were halved while [B] remained constant.

Answer: ___________________________________

Lesson 2-8 — Factors Affecting the Rate of Reaction

✨ Engaging Introduction

A piece of iron left outdoors rusts slowly over months, while a sparkler burns in seconds. A hydrogen peroxide solution may remain stable in a bottle for weeks, yet it bubbles vigorously when poured onto a wound. Why do some reactions occur almost instantly while others take years?

In this lesson, we will investigate how factors such as concentration, temperature, catalysts, surface area, and light influence the speed of chemical reactions and how scientists use evidence to explain these changes.

> ❓ **Guiding question**: How can we control the speed of a chemical reaction to make processes safer, faster, and more efficient?

Inquiry-Based Investigation

Materials

- Effervescent tablets
- Three beakers
- Water at different temperatures (cold, room temperature, hot)
- Stopwatch

Procedure

1. Place equal volumes of water in three beakers.
2. Adjust temperatures to cold, room temperature, and hot.
3. Add identical tablets simultaneously.
4. Measure the time required for each tablet to dissolve completely.

What is your observation?

……

What is your conclusion?

……

⭐ **structure of science**

Key Fact :

A chemical reaction occurs only when particles collide with enough energy to overcome the activation energy barrier.

Key Concepts:

· Catalyst
· Homogeneous catalyst
· Heterogeneous catalyst
· Enzyme
· Transition state
· Activation energy

Point!

1. Factors Affecting the Rate of Reaction

(a) Chemical reactions and collisions of particles

Chemical reactions occur when particles such as atoms, molecules, and ions that make up a substance collide with each other. An increase in the number of particle collisions increases the rate of reaction.

(b) Conditions for changing the rate of reaction

(i) **Concentration:** The higher the concentration of reactants, the greater the rate of reaction.

(ii) **Temperature:** As the temperature increases, the rate of reaction increases significantly.

(iii) **Catalyst:** A substance that increases the rate of reaction without itself undergoing any permanent chemical change.

- **Homogeneous catalyst:** A catalyst that exists in the same phase as the reactants.
- **Heterogeneous catalyst:** A catalyst that exists in a different phase from the reactants.
- **Enzyme:** A catalyst that works within a living organism. Composed of proteins.

(iv) Surface area of solid: Increasing the surface area of a solid increases the rate of reaction.

> **Example:** When the same amount of hydrochloric acid is added to magnesium ribbon and powdered magnesium, hydrogen is generated more vigorously from the powdered magnesium.

(v) Photochemical reaction: Chemical reactions may occur when a substance absorbs light energy.

Problem-Solving

A chemical company wants to increase the production rate of a product without increasing the amount of reactants.

Question

Which modification would most effectively increase the reaction rate?

A. Lower the temperature

B. Add a suitable catalyst

C. Reduce the surface area of reactants

D. Dilute the reactants

Reasoning

write your Reasoning .

....

2. Activation energy

(a) **Transition state:** A high-energy state where the rearrangement of atoms occurs. When a chemical reaction proceeds, colliding particles pass through this state.

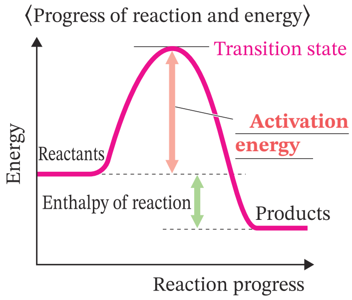


Figure 2.8.1 Activation Energy

(b) **Activation energy:** The minimum energy required to bring reactants to the transition state (Figure on the right 2.8.1).

- For a chemical reaction to proceed, particles must collide with energy equal to or greater than the activation energy.

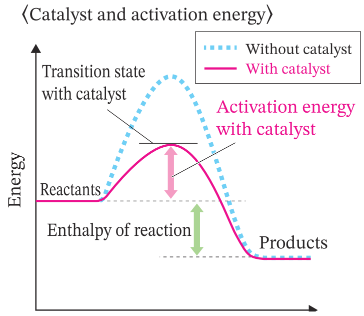


Figure 2.8.2 Catalyst and Activation Energy

(c) **Catalyst and activation energy Figure on the right 2.8.2):**

Using a catalyst increases the rate of reaction.

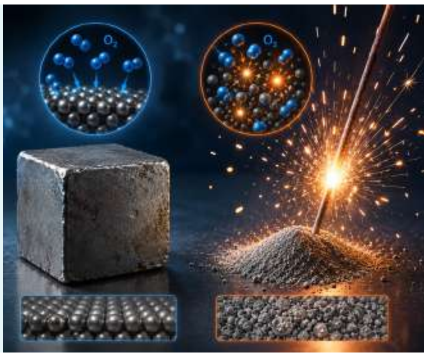
⟨Reason⟩ This is because the catalyst provides an alternative reaction pathway with a lower activation energy, making it easier for reactants to reach the transition state.

Reference table — factors that change the rate

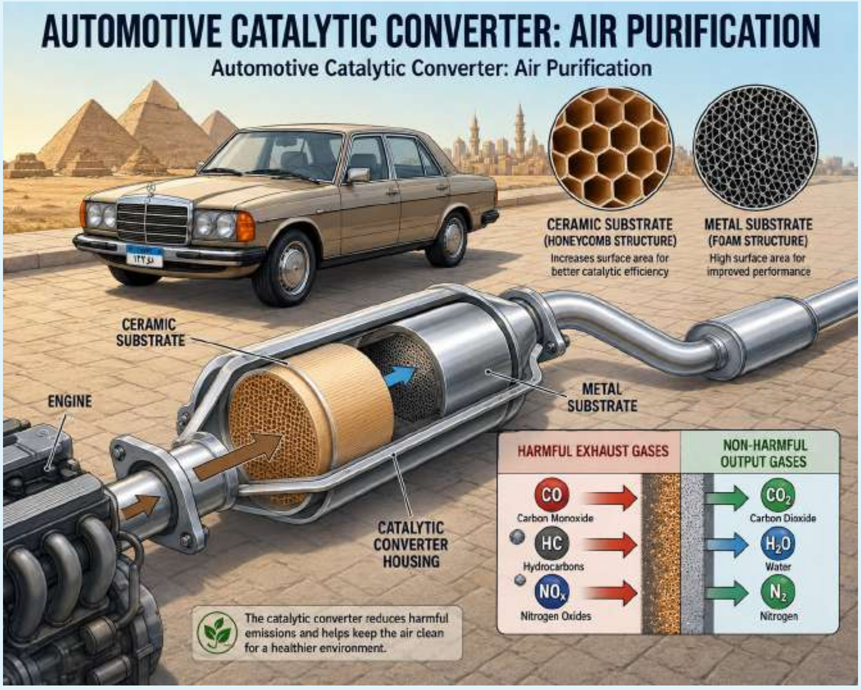
Factor	Effect on rate	Why
Concentration ↑	rate increases	more frequent particle collisions
Temperature ↑	rate increases significantly	more collisions have energy ≥ activation energy
Catalyst	rate increases	provides an alternative path with lower activation energy
Surface area ↑	rate increases	more contact area / more collisions
Light (photochemical)	can start a reaction	substance absorbs light energy

Life Application

When a metal such as iron or aluminum is present as a large solid piece, it reacts relatively slowly with oxygen in the air. However, when the same metal is ground into a fine powder and used in fireworks or sparklers, it burns rapidly and produces bright light, sparks, and heat.



Life Application



Everyday Catalysts: The Three-Way Catalyst in Automobile Exhaust Purification Harmful substances such as carbon monoxide (CO), hydrocarbons, and nitrogen oxides (NO_x) are contained in car exhaust gas. To rapidly convert these into harmless products — carbon dioxide (CO_2), water (H_2O), and nitrogen (N_2) — a **Three-Way Catalyst is** used.

(i) Increasing the Rate of Reaction with a Catalyst Although harmful substances in exhaust gas can naturally react with oxygen to become harmless without a catalyst, the rate of reaction is extremely slow — insufficient to process the continuous stream of gases emitted while driving. The Three-Way Catalyst, which uses noble metals like platinum (Pt) and palladium (Pd), dramatically lowers the required **activation energy**, so the reactions are completed almost instantly and the exhaust gas is purified.

(ii) Why the Honeycomb Structure?

The actual Three-Way Catalyst has a fine, ceramic mesh structure, known as a **Honeycomb Structure** (resembling a beehive), used to maximise the **surface area** of contact between the catalyst and the gases. Since the rate of reaction is proportional to the contact area between the reactants and the catalyst, this structure allows efficient purification of large volumes of exhaust gas within a limited space.

Warm Up

Answer the following questions.

(a) The following (i) to (iv) describe factors affecting the rate of reaction. For each, choose the most closely related factor from A to E below.

(i) Hydrogen peroxide solution is stored at a low temperature.

(ii) When dilute hydrogen peroxide solution is applied to the skin, no particular change is observed, but when applied to a wound, it foams vigorously.

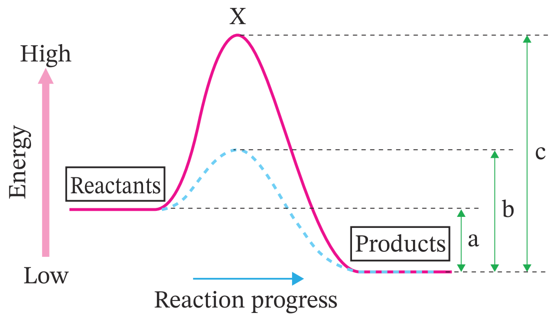
(iii) When a small piece of zinc is placed in 1 mol/L hydrochloric acid and aqueous acetic acid solution respectively, hydrogen gas is produced more rapidly in hydrochloric acid.

(iv) Nitric acid is stored in an amber bottle.

A. Concentration　　B. Temperature

C. Light　　D. Catalyst　　E. Pressure

(b) The figure on the right shows the relationship between the reaction pathway and energy for a certain reaction with and without a catalyst. **Answer the following questions.**



(i) Is this reaction an exothermic reaction or an endothermic reaction?

(ii) What is state X called?

(iii) Express the following values using a to c in the figure.

A. Activation energy without a catalyst

B. Activation energy with a catalyst

C. Enthalpy of reaction with a catalyst

Explanation

(a)(i) Since the rate of reaction increases when the temperature is high, it is kept at a low temperature to make the reaction less likely to occur. **B**

(ii) When applied to a wound, enzymes present in living tissues act as catalysts, significantly increasing the rate of reaction. **D**

(iii) Since hydrogen chloride is a strong acid and acetic acid is a weak acid, the hydrogen ion concentration $[H^+]$ is greater in hydrochloric acid, resulting in a greater rate of formation of hydrogen. **A**

(iv) Since nitric acid is easily decomposed by light, it is stored in a light-shielding amber bottle. **C**

(b)(i) From the figure, the energy decreases due to the reaction. A decrease in energy indicates an exothermic reaction, which releases energy during the reaction.

(ii) Transition state.

(iii) The solid line with the higher activation energy represents the reaction without a catalyst, and the dashed line

represents the reaction with a catalyst.

A: The activation energy is the difference between the energy possessed by the reactants and the energy at the

transition state. Therefore, c−a

B: b−a

C: The enthalpy of reaction is the difference in energy between the reactants and the products.

Also, since the energy possessed by the products is smaller than the energy possessed by the reactants, the sign is negative. −a .

Try

1. Two identical reactions are performed. In Experiment A, powdered calcium carbonate is used; in Experiment B, large marble chips are used. Which statement best explains why Experiment A is faster?

A. Powder has lower activation energy.

B. Powder has greater surface area.

C. Powder increases temperature.

D. Powder acts as a catalyst.

Answer: ___________________________________

2. A reaction becomes four times faster when the temperature is increased. Which explanation is most accurate?

A. More particles are created.

B. More collisions have sufficient energy to react.

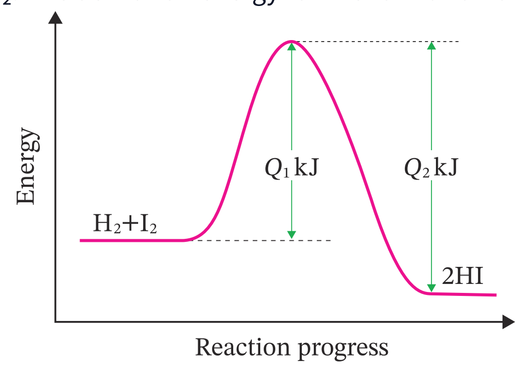
C. Activation energy increases.

D. Reactants disappear instantly.

Answer: ___________________________________

3. Read the following text and answer the questions below.

The figure on the right shows the energy change of the reaction in which hydrogen iodide HI is formed from hydrogen H_2 and iodine I_2. The activation energy for the formation of hydrogen iodide is (**A**) kJ, the enthalpy of reaction is (**B**) kJ, the activation energy for the decomposition of hydrogen iodide is (**C**) kJ, and the enthalpy of reaction is (**D**) kJ. When a catalyst is added to increase the rate of reaction, the activation energy of the formation reaction (**E**). At this time, the enthalpy of reaction for the formation (**F**).



(a) Choose appropriate symbols or words for A to F in the text from (i) to (vii) below.

(i) Q_1	(ii) Q_2	(iii) $Q_1−Q_2$	(iv) $Q_2−Q_1$
(v) Increases	(vi) Decreases	(vii) Does not change	

(b) As stated in the underlined part, adding a catalyst increases the rate of reaction. Explain the reason for this.

Answer: ___________________________________

Engineering Design Challenge

Design a simple investigation to determine which factor has the greatest effect on the rate of dissolving an effervescent tablet:

• Temperature

• Surface area

• Concentration

Students must:

1. Form a hypothesis.

Answer: ___________________________________

2. Identify variables.

Answer: ___________________________________

3. Collect data.

Answer: ___________________________________

4. Put conclusions using evidence.

Answer: ___________________________________

Exercise

1. For each statement regarding the rate of reaction, choose the factor from the word bank below that is most closely related to the statement. The same factor may be chosen multiple times.

(a) Wood burns more rapidly when it is split into small pieces.

(b) Concentrated nitric acid is stored in an amber bottle.

(c) When a small amount of aqueous iron(III) chloride solution is added to hydrogen peroxide solution, oxygen is generated rapidly.

(d) When equal masses of zinc are added to 10 mL each of 1 mol/L hydrochloric acid and acetic acid, hydrogen is generated more rapidly in the hydrochloric acid.

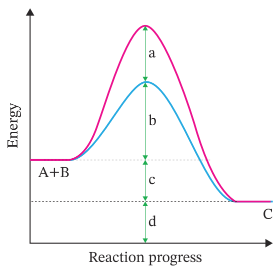
(e) An incense stick burns faster in oxygen than in air.

(f) Hydrogen peroxide is better stored in a refrigerator if possible.

[Word bank: Temperature, Concentration, Pressure, Surface area, Catalyst, Light]

Answer: ___________________________________

2. The figure shows the change in the transition state for the reaction $A + B \longrightarrow C$ with and without a catalyst. Answer the following questions.



(a)Express the following energies using the symbols a to d in the figure.

(i)Activation energy of $A + B \longrightarrow C$ without a catalyst

(ii)Activation energy of $A + B \longrightarrow C$ with a catalyst

(iii)Activation energy when the reaction $C \longrightarrow A + B$ occurs without a catalyst

(iv)Enthalpy of reaction when the reaction $A + B \longrightarrow C$ occurs

(b)Is the reaction $A + B \longrightarrow C$ an exothermic reaction or an endothermic reaction?

Answer: ___________________________________

Lesson 2-9 — Chemical Equilibrium

✨ Engaging Introduction

Imagine opening a bottle of carbonated soft drink. As soon as the bottle is opened, bubbles begin to escape. If the bottle remains closed, carbon dioxide stays dissolved in the drink. Why does the system behave differently when conditions change?

Similarly, in our bodies, oxygen continuously binds to and releases from hemoglobin, allowing oxygen transport to tissues. These processes involve reactions that proceed in both directions and eventually reach a balance called **chemical equilibrium**.

In this lesson, we will explore how reversible reactions reach equilibrium and how chemists use equilibrium concepts to predict and control chemical systems.

> ❓ **Guiding question:** How can a chemical reaction appear to stop while reactants and products continue to react?

🔭 Exploration

A student places a sample of brown nitrogen dioxide gas (NO_2) in a sealed transparent syringe . At first, the gas appears dark brown. After some time,

What is your observation?

……………………………………………………………………………

when the gas in compressed?

……………………………………………………………………………

What is your Prediction?

If the temperature changes, the equilibrium position may shift, changing the color intensity again.?

……………………………………………………………………………

📌 **Point!**

1. Reversible and irreversible Reactions

(a) **Reversible reaction:** A reaction that proceeds in both directions. Expressed using a double arrow "$\rightleftarrows$"

> **Example:** When hydrogen and iodine are heated, hydrogen iodide is formed. Conversely, when hydrogen iodide is heated, it decomposes into hydrogen and iodine.
> $H_2 + I_2 \rightleftharpoons 2HI$

• In a reversible reaction, the reaction proceeding to the right is called the forward reaction, and the reaction proceeding to the left is called the backward reaction.

(b) **Irreversible reaction:** A reaction that proceeds only in one direction. Expressed using a single arrow "$\longrightarrow$".

⭐ **structure of science**

Key Fact :

At chemical equilibrium, the forward and reverse reactions continue to occur, but at equal rates, causing concentrations to remain constant.

Key Concepts:

· Reversible reaction
· Irreversible reaction
· Dynamic equilibrium
· Equilibrium constant (K)
· Forward reaction
· Reverse reaction.

Example: When steel wool is heated, iron(II) oxide FeO is formed, but the reverse reaction does not occur. $2Fe + O_2 \rightarrow 2FeO$

2. Chemical Equilibrium

(a) **State of chemical equilibrium (Equilibrium state):** A state where the rates of the forward and reverse reactions become equal, and the concentrations of reactants and products remain constant.

(b) **Law of chemical equilibrium (Law of mass action)**

• For reactants A, B and products C, D, when the reversible reaction $aA + bB \rightleftharpoons cC + dD$ is in an equilibrium state, the following relationship holds among the concentrations [A], [B], [C], and [D] of each substance.

$$\frac{[C]^c[D]^d}{[A]^a[B]^b} = K \text{ (constant)}$$

Note: K in this equation is called the **equilibrium constant** of chemical equilibrium, and it is constant as long as the temperature does not change.

• The equilibrium constant in reactions between solids and gases is expressed only by the concentration of the gases.

This is because the concentrations of pure solid substances remain constant and do not affect the equilibrium expression.

Example: The equilibrium constant K in the reaction $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$ is as follows.

$$K = \frac{[CO]^2}{[CO_2]}$$

• The unit of the equilibrium constant is determined by calculating the units together.

Example: The unit of $K = \frac{[B]}{[A]^3}$ is $[(mol/L)^{-2}]$.

(c)Calculating the Equilibrium Constant

1. Write the balanced chemical equation and record the initial amount of substance (or concentration) of each reactant and product.
2. Determine the change in the amount of substance (or concentration) of each substance as the reaction proceeds toward equilibrium.
3. Calculate the amount of substance (or concentration) of each substance at equilibrium.
4. Substitute the equilibrium concentrations into the equilibrium constant expression and calculate the equilibrium constant.

	$N_2O_4 \rightleftharpoons$	$2NO_2$	
1 Initial	0.50	0	[mol]
2 Change	−0.30	+0.60	[mol]
3 Equilibrium	0.20	0.60	[mol]

4 Equilibrium constant $K = \frac{\left(\frac{0.60\ mol}{10\ L}\right)^2}{\left(\frac{0.20\ mol}{10\ L}\right)}$

Think Critically

Analyze Experimental Data

Scientists often increase the pressure in the Haber Process to produce more ammonia.

If increasing pressure requires more energy and higher costs, how should engineers decide whether the increase in ammonia production is worth the extra cost?

....

....

....

Warm Up

When 5.50 mol of hydrogen and 4.00 mol of iodine are placed in a 100 L container and kept at a certain temperature, the reaction below occurs, reaching an equilibrium state. At this time, 7.00 mol of hydrogen iodide was formed. Answer the following questions.

$H_2 + I_2 \rightleftharpoons 2HI$

(a) Calculate the equilibrium constant for this reaction to two significant figures.

(b) When 3.00 mol of hydrogen and 3.00 mol of iodine are placed in the same container and kept at the same temperature as in (a), how many moles of hydrogen iodide will be formed? Give your answer to two significant figures.

> **Be Global Citizen**
>
> Increasing atmospheric carbon dioxide dissolves in oceans and affects the equilibrium of carbonate ions needed by corals and shellfish.

Explanation

(a) It is helpful to write the balanced chemical equation and construct a reaction table.

	H_2	+ I_2	$\rightleftharpoons$ $2HI$	
Initial	5.50	4.00	0	[mol]
Change	−3.50	−3.50	+7.00	[mol]
Equilibrium	2.00	0.50	7.00	[mol]

The equilibrium constant $K = \frac{[HI]^2}{[H_2][I_2]}$, and since the size of the container is 100 L, the equilibrium constant of this reaction is,

$$\frac{\left(\frac{7.00\ \text{mol}}{100\ \text{L}}\right)^2}{\left(\frac{2.00\ \text{mol}}{100\ \text{L}}\right) \times \left(\frac{0.50\ \text{mol}}{100\ \text{L}}\right)} = \underline{49}$$

(b) Assuming that x mol each of H_2 and I_2 react to reach equilibrium, the amounts of substance are as follows.

	H_2	+	I_2	$\rightleftarrows$	2HI	
Initial	3.00		3.00		0	[mol]
Change	−x		−x		+2x	[mol]
Equilibrium	3.00−x		3.00−x		2x	[mol]

Since the equilibrium constant is constant at the same temperature, from (a), the equilibrium constant is 49. Therefore,

$$\frac{\left(\frac{2x\,[\text{mol}]}{100\text{ L}}\right)^2}{\left(\frac{3.00-x\,[\text{mol}]}{100\text{ L}}\right)\times\left(\frac{3.00-x\,[\text{mol}]}{100\text{ L}}\right)} = 49$$

$$\frac{(2x)^2}{(3.00-x)^2} = 49$$

$$\frac{2x}{3.00-x} = \pm 7$$

$$x = 2.33\text{ mol}, 4.2\text{ mol}$$

Since the initial amounts of hydrogen and iodine are 3.00 mol, x must be less than 3.00 mol ($0 < x < 3.00$). Therefore,

$x = 2.33$ mol

Therefore, the hydrogen iodide formed at equilibrium is $2.33\times2 = 4.66 \approx 4.7$ mol 4.7 mol

Try

Solve the Following Questions

1. A mixture of 1.00 mol of acetic acid and 1.00 mol of ethanol reacts at a constant temperature. At equilibrium, 0.25 mol of acetic acid remains. Give your answers to two significant figures.

$$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$$

(a) Calculate the equilibrium constant for this reaction.

Answer: ______________________________________

(b) When a mixture of 1.00 mol of acetic acid, 1.00 mol of ethanol, and 4.00 mol of water is reacted and reaches equilibrium at the same temperature, how many moles of ethyl acetate $CH_3COOC_2H_5$ are formed?

Answer: ______________________________________

2. A mixture of 1.00 mol of H_2 and 1.00 mol of I_2 is placed in a container of constant volume at a constant temperature. At equilibrium, 1.6 mol of HI is present. Give your answers to two significant figures.

(a) Calculate the equilibrium constant for this reaction.

Answer: ______________________________________

(b) When 1.6 mol each of H_2 and I_2 are placed in a container of constant volume at the same temperature, determine the amount of substance of HI at equilibrium.

Answer: ______________________________________

Exercise

Choose the correct answer:

1. For the reaction $N_2O_4 \rightleftharpoons 2NO_2$, the equilibrium constant K is correctly written as:

A. $K = [N_2O_4]/[NO_2]^2$

B. $K = [NO_2]^2/[N_2O_4]$

C. $K = [NO_2]/[N_2O_4]$

D. $K = [N_2O_4]/[NO_2]$

Answer: ___________________________________

2. A soft drink bottle is opened and bubbles rapidly form. Which explanation is most accurate?

A. Equilibrium shifts as pressure decreases.

B. Water becomes a catalyst.

C. The reaction becomes irreversible.

D. The temperature increases

Answer: ___________________________________

Answer on the Following Questions

3. For the following reactions, write the equilibrium constant expression in terms of molar concentrations. Also, state the units of the equilibrium constant if applicable. Assume all substances are gases unless otherwise specified.

(a) $N_2 + O_2 \rightleftharpoons 2NO$

(b) $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

(c) $CO_2 + H_2 \rightleftharpoons CO + H_2O$

(d) $C_{(s)} + H_2O \rightleftharpoons CO + H_2$

4. When acetic acid CH_3COOH and ethanol C_2H_5OH react in the presence of concentrated sulfuric acid as a catalyst, ethyl acetate $CH_3COOC_2H_5$ and water are formed, and the system reaches equilibrium. Answer the questions below to two significant figures.

$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$

(a) When 5.0 mol of acetic acid and 8.0 mol of ethanol are mixed and reacted, the equilibrium state is reached, and 4.0 mol of ethyl acetate is formed. Calculate the equilibrium constant, K, for this reaction.

Answer: ___________________________________

(b) Under the same temperature as (a), when 8.0 mol of acetic acid and 8.0 mol of ethanol are mixed and reacted, how many moles of ethyl acetate are formed?

Answer: ___________________________________

Lesson 2-10 — Le Chatelier's Principle

✨ Engaging Introduction

Ammonia is one of the world's most important chemicals because it is used to manufacture fertilizers that help feed billions of people.

Scientists discovered that increasing pressure in the reaction:

$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

produces more ammonia.

If the pressure inside the reaction vessel is suddenly increased, what do you predict will happen to the amount of ammonia produced? Explain your reasoning.

> ❓ **Guiding question:** How do changes in concentration, pressure, temperature, and catalysts affect chemical equilibrium?

🌍 Exploration

A fertilizer factory produces ammonia using nitrogen and hydrogen gases. Engineers notice that increasing pressure inside the reactor dramatically increases ammonia production, while increasing temperature reduces the amount produced.

Predict

Before studying Le Chatelier's Principle:

What do you think will happen to ammonia production if:

1. More nitrogen gas is added?

....

2. The pressure is doubled?

....

3. The temperature is increased?

....

📌 **Point!**

1. Direction of equilibrium shift

Le Chatelier's principle:

When a reversible reaction is in an equilibrium state, if conditions such as concentration, pressure, and temperature are changed, the equilibrium shifts in a direction that counteracts the effect, reaching a new equilibrium state.

(a) Effect of Concentration:

When the concentration of a certain component increases, the equilibrium shifts in the direction that **decreases** its concentration.

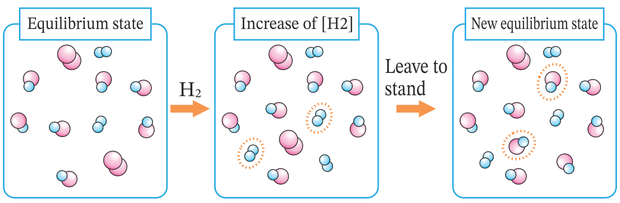


Figure2.10.1 the effect of Concentration

⭐ **structure of science**

Key Fact :

A chemical equilibrium shifts in the direction that opposes any change imposed on the system.

Key Concepts:

- Reversible reactions
- Chemical equilibrium
- Equilibrium shift
- Concentration
- Pressure
- Temperature
- Catalyst

Key law

— Le Chatelier's Principle.

Example: For the equilibrium system $H_2 + I_2 \rightleftharpoons 2HI$, if $[H_2]$ increases under a constant temperature, it will eventually reach a new equilibrium state by decreasing $[H_2]$. ⇒ The equilibrium shifts to the **right**. (Figure 2.10.1)

(b) Effect of Pressure:

When the pressure of a mixed gas increases, the equilibrium shifts in the direction where the total number of gas molecules **decreases**. (Figure 2.10.2)

In a reaction where the total number of gas molecules does not change before and after the reaction, the equilibrium **does not shift**.

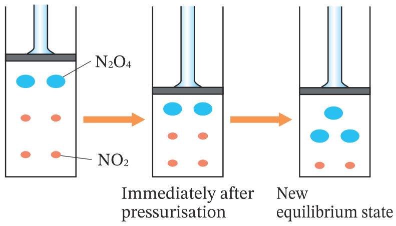


Figure 2.10.2 the effect of Pressure

(c) Effect of Temperature:

When the temperature is lowered, the equilibrium shifts in the direction of the **exothermic reaction**, and when the temperature is raised, the equilibrium shifts in the direction of the **endothermic reaction**.

(d) Effect of a Catalyst: Using a catalyst shortens the time required to reach the equilibrium state, but the equilibrium **does not shift**.

(e) Effect of an Inert Gas (Figure 2.10.3)

When the reversible reaction $2NO_2 \rightleftharpoons N_2O_4$ is in an equilibrium state,

(i) Argon Ar is added at a constant temperature and volume When the volume is constant, the concentration and partial pressure of the substances involved in the reaction do not change.
⇒ The equilibrium does not shift .

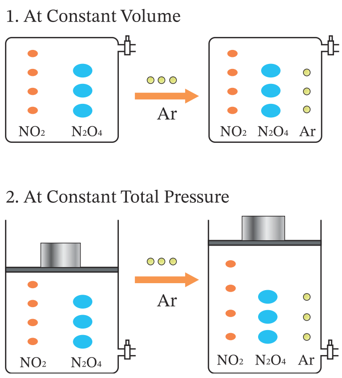


Figure 2.10.3
The effect of an Inert Gas

(ii) Argon Ar is added at a constant temperature and total pressure When the total pressure is constant, the volume increases by the amount of Ar added, so the partial pressure of the gases involved in the reaction inside the container decreases.

⇒ The equilibrium shifts to the left .

2. Synthesis of ammonia

The reaction to synthesize ammonia,
$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$ $\Delta H = -92$ kJ,
is an exothermic reaction in which the total number of molecules decreases .

⇒ By making the pressure high and the temperature low , the equilibrium shifts to the right, and the yield of ammonia increases. (Figure 2.10.4).

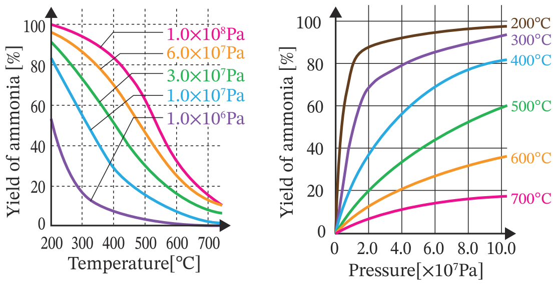


Figure 2.10.4 Synthesis of ammonia

Life Application

Ammonia Production (Haber Process)

The global production of fertilizers depends on controlling equilibrium conditions to maximize ammonia yield. Without equilibrium control, modern agriculture could not produce enough food for the world's growing population.

Warm Up

Answer the following questions.

(a) For each of the following systems at equilibrium, in which direction will the equilibrium shift when subjected to the indicated change? Answer with the symbols (A) Left, (B) Right, (C) Does not shift.

(i) $3O_2 \rightleftharpoons 2O_3$	(Adding oxygen)
(ii) $2SO_2 + O_2 \rightleftharpoons 2SO_3$	(Adding a catalyst)
(iii) $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$	(Adding ammonium chloride)

(b) In the synthesis of ammonia (an exothermic reaction), a low temperature is favorable for shifting the equilibrium position to the right. However, the actual industrial manufacturing process is carried out at a high temperature of approximately 400–500°C. Explain the reason for this in terms of the rate of reaction.

Explanation

(a) **(i)** When oxygen is added, the concentration of oxygen increases, so the equilibrium shifts in the direction that decreases the concentration of oxygen. Therefore, the equilibrium shifts to the right. **(B)**

(ii) Even if a catalyst is added, the equilibrium does not shift. **(C)**

(iii) When ammonium chloride is added, it ionizes in the aqueous solution, increasing the concentration of ammonium ions. The equilibrium then shifts in the direction that consumes the added ammonium ions. The equilibrium shifts to the left. **(A)**

(b) If the temperature is too low, the rate of reaction is too small, and the product cannot be obtained within a practical timeframe.

NOS: Identifying variables

A student wants to investigate the effect of temperature on ammonia equilibrium.

1. What is the independent variable? ....

2. What is the dependent variable? ....

3. Which variables should be controlled to ensure a fair test?....................................

Try

Answer on the Following Questions

1. For each of the following equilibrium reactions, determine the direction of the shift in equilibrium position when the condition change shown in ( ) is applied. Answer with "Right", "Left", or „Does not shift".

(a) $H^+ + OH^- \rightleftarrows H_2O_{(l)}$ $\Delta H = -56$ kJ (Cooling)

(b) $CO + 2H_2 \rightleftarrows CH_3OH_{(g)}$ (Increasing the pressure)

(c) $CO + 2H_2 \rightleftarrows CH_3OH_{(g)}$ (Adding a catalyst)

(d) $N_2 + 3H_2 \rightleftarrows 2NH_3$ (Adding Ar at constant volume)

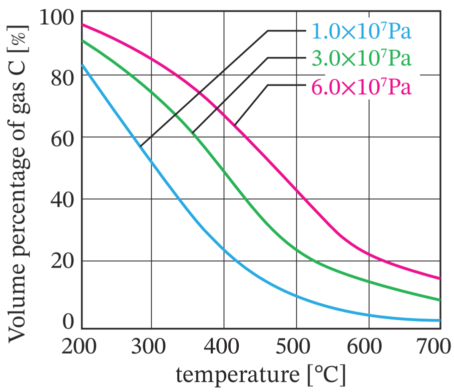
(e) $N_2 + 3H_2 \rightleftarrows 2NH_3$ (Adding Ar at constant pressure)

(f) $AgCl_{(s)} \rightleftarrows Ag^+ + Cl^-$ (Adding $HCl_{(g)}$)

2. Consider the following chemical equilibrium involving gases A, B, and C:

$aA + bB \rightleftarrows cC$ (a, b, and c are coefficients)

The relationship between the volume percentage of C and temperature in the equilibrium state of this reaction is as shown in the figure on the right. Answer the following questions.



(a) Which of the following relationships is correct?

A a+b<c **B** a+b>c **C** a+b=c

Answer: ________________________________

(b) Is the forward reaction of this reaction an exothermic or endothermic reaction?

Answer: ________________________________

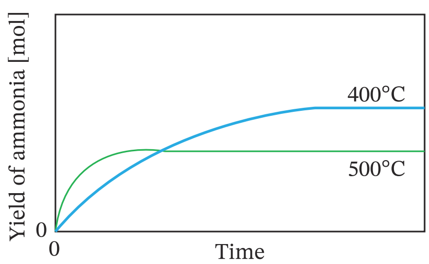
(c) As the temperature is raised, how does the equilibrium constant K of this reaction change?

Answer: ________________________________

3. Ammonia is synthesized from nitrogen and hydrogen by the following reaction.

$N_2 + 3H_2 \rightleftarrows 2NH_3$

When 1 mol of nitrogen and 3 mol of hydrogen react at a constant pressure in the presence of an iron-based catalyst, the amount of ammonia produced increases until equilibrium is reached. The change in the amount of ammonia formed over time at 400 °C and 500 °C is shown by solid lines in the figure. With reference to this figure, choose one correct statement from the following.



A. The formation reaction of ammonia is an endothermic reaction.

B. The equilibrium constant of the above chemical equation at 500°C is smaller than the value at 400°C.

C. The rate at which ammonia is formed increases over time at both 400°C and 500°C.

D. When experimenting at a higher pressure, the rate of reaction increases, but the amount of ammonia formed at equilibrium decreases.

Exercise

Answer on the Following Questions

1. For each of the following equilibrium reactions, determine the direction of the shift when the condition change shown in ( ) is applied. Answer with „Right”, „Left”, or „Does not shift”. *Note: (g) = gas, (l) = liquid, (s) = solid, (aq) = aqueous solution; ΔH = enthalpy change.*

(a) $2HI \rightleftarrows H_2 + I_{2(g)}$ (Adding HI)

(b) $2NO_2 \rightleftarrows N_2O_4$ (Removing NO_2)

(c) $2HI \rightleftarrows H_2 + I_{2(g)}$ $\Delta H = 9$ kJ (Raising the temperature)

(d) $2NO_2 \rightleftarrows N_2O_4$ $\Delta H = -57$ kJ (Lowering the temperature)

(e) $2HI \rightleftarrows H_2 + I_{2(g)}$ (Increasing the pressure)

(f) $2SO_2 + O_2 \rightleftarrows 2SO_3$ (Adding Ar at constant pressure)

(g) $C_{(s)} + H_2O_{(g)} \rightleftarrows CO + H_2$ (Increasing the pressure)

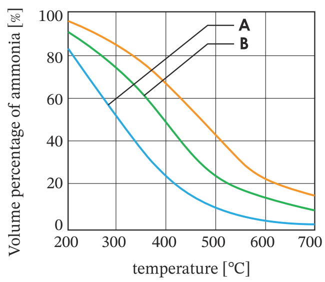
(h) $2SO_2 + O_2 \rightleftarrows 2SO_3$ (Adding a catalyst)

Answer: ___________________________________

2.The following chemical equation is the formation reaction of ammonia.

$N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)}$

When a mixed gas of nitrogen and hydrogen at a ratio of 1 : 3 was reacted under 1.0×10^7 Pa, 3.0×10^7 Pa, and 6.0×10^7 Pa at various temperatures, the results shown in the right figure were obtained. Answer the following questions.



(a) Determine the pressure in Pa represented by curves A and B in the figure, respectively.

Answer: ___________________________________

(b) Based on the figure, is the formation of ammonia exothermic or endothermic?

Answer: ___________________________________

Lesson 2-11 — Ionisation Equilibrium

✨ Engaging Introduction

Imagine opening two bottles of clear liquid. Both look like pure water, yet when you insert an electrical conductivity tester, one solution makes the bulb glow brightly while the other produces only a faint light. In another case, a third solution shows no electrical response at all, even though it completely dissolves in water.

In this lesson, we will explore how chemists use ion formation and equilibrium concepts to explain why solutions conduct electricity differently and how these invisible processes determine the behavior of matter in water.

> ❓ **Guiding question:** How can scientists explain and predict the electrical behavior of solutions based on the extent of ion formation in water?

Exploration

Strong, Weak, or Non-Electrolyte?

Consider the following substances dissolved in water:

- Hydrochloric acid (HCl)
- Acetic acid (CH_3COOH)
- Glucose ($C_6H_{12}O_6$)

Without studying the lesson:

- Which substance will produce the highest number of ions? ....
- Which substance will not produce ions? ....
- Which will produce a weak electrical current? ....

Record your initial ideas and revisit them after studying electrolytes and ionization equilibrium.

....

📌 **Point!**

1. The Nature of Electrolytes: Strong and Weak Electrolytes

When a substance dissolves in water, the phenomenon where it separates into ions is called **dissociation** (or **ionisation**). Substances are broadly classified into two groups based on their extent of dissociation, known as the **degree of dissociation** (alpha, α).

- **Strong Electrolytes** (Strong Acids and Strong Bases): substances like hydrogen chloride (HCl) and sodium hydroxide ($NaOH$) that dissociate almost completely into ions in aqueous solution. For calculation purposes, $\alpha \approx 1$.
- **Weak Electrolytes** (Weak Acids and Weak Bases): substances like acetic acid (CH_3COOH) and ammonia (NH_3). Only a fraction of their molecules turn into ions, while the majority remain as molecules.

⭐ **structure of science**

Key Fact :

Strong electrolytes dissociate completely into ions in aqueous solutions. • Weak electrolytes partially dissociate and establish equilibrium in water. • Electrical conductivity depends on the number of free ions in solution.

Key Concepts:

· Dissociation
· Ionization
· Electrolyte
· Strong electrolyte
· Weak electrolyte
· Degree of dissociation (α)
· Dynamic equilibrium
· Dissociation constant (K_a)
· Ionic conductivity
· Ion mobility
· Solvation

2. Dissociation Equilibrium as a Reversible Reaction

When a weak electrolyte is dissolved in water, two simultaneous reactions occur: the **forward reaction** (molecules forming ions) and the **reverse reaction** (ions reforming molecules). Eventually, the rates of both reactions become equal, and the apparent change stops. This condition is called **dissociation equilibrium**.

For example, the dissociation equilibrium of acetic acid can be described as follows:

$CH_3COOH \leftrightarrows CH_3COO^- + H^+$

In this state of equilibrium, provided the temperature is constant, the following Dissociation Constant (K_a) is established.

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

Note: K_a is a constant that depends only on temperature and does not change with concentration.-----

Let us mathematically prove why the degree of dissociation (α) increases when the concentration is diluted:

Substance	Initial Conc.	Change in Conc.	Equilibrium Conc.
CH_3COOH	c	−cα	c (1−α)
CH_3COO^-	0	+cα	cα
H^+	0	+cα	cα

Substituting these into the K_a expression:

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} = \frac{c\alpha \times c\alpha}{c\,(1-\alpha)} = \frac{c\alpha^2}{1-\alpha}$$

Since α is extremely small compared to 1 for a weak acid (α << 1), we can make the approximation that 1−α ≈ 1.

Thus, $K_a \approx c\alpha^2$, which rearranges to $\alpha = \sqrt{\frac{K_a}{c}}$

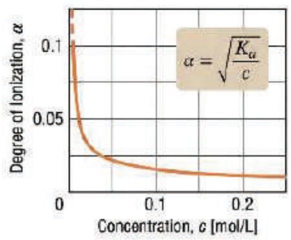


figure 2.11.1 Relationship between Acetic Acid Concentration and Degree of Ionization (calculated values)

This equation demonstrates that the smaller the concentration c (i.e., the more dilute the solution), the larger the degree of dissociation α becomes.

Think Critically

A student claims:

"Acetic acid is a weak acid because it does not dissolve well in water."

Claim–Evidence–Reasoning

Students evaluate:

- Is the claim scientifically valid?..
- What is the difference between solubility and ionization? ....
- Can a substance be highly soluble but still weakly ionized? Explain..............................

3. The Ionic Product of Water (Kw) and its Relation to Temperature

Even pure water dissociates to a very small extent.

$H_2O \leftrightarrows H^+ + OH^-$

The Ionic Product of Water (K_w) is derived from this equilibrium.

$K_w = [H^+][OH^-] = 1.0\times10^{-14}\ (mol/L)^2$ (at 25°C)

It is important to note that the dissociation of water is an endothermic reaction. According to Le Chatelier's Principle, increasing the temperature shifts the equilibrium to the right, which is the direction that promotes dissociation. This means that as the temperature rises, the value of K_w increases. Consequently, even a neutral solution will have a $[H^+]$ concentration greater than 1.0×10^{-7} mol/L, resulting in a pH value less than 7..

4. Definition of pH

PH is a concise way to express the strength of acidity or basicity using base−10 logarithms.

pH = $-\log_{10}[H^+]$

Note:

For a basic (alkaline) solution, the usual procedure is to first calculate $[OH^-]$ and then use the relationship $K_w = [H^+][OH^-]$ to find $[H^+]$. Alternatively, you may calculate pOH = $-\log_{10}[OH^-]$ and then use the relationship pH + pOH = 14 (at 25°C).

Life Application

Ionization equilibrium in Real Life

The concept of electrolytes is essential in:

- Human nerve signal transmission (ionic movement in cells)
- Medical saline solutions (controlled ion concentration)
- Batteries and electrochemical cells
- Industrial electrolysis processes

Weak and strong electrolytes determine how efficiently charge moves in biological and industrial systems

Warm Up

Answer the following questions:

(a) Calculate the pH of a 0.10 mol/L aqueous acetic acid solution (degree of dissociation 0.016) to one decimal place. ($\log_{10} 1.6 = 0.2$)

(b) Consider a 0.020 M aqueous solution of ammonia at 25°C. Assume the dissociation constant of ammonia is $K = 1.8\times10^{-5}$ mol/L and the ionic product of water is $Kw = 1.0\times10^{-14}$ mol^2/L^2. ($\log_{10} 2.0 = 0.30$, $\log_{10} 3.0 = 0.48$)

(i) Calculate the degree of dissociation of the aqueous ammonia.

(ii) Calculate the hydroxide ion concentration.

(iii) Calculate the pH to one decimal place.

Explanation

(a) Since pH = − $\log_{10}[H^+]$, first find the hydrogen ion concentration $[H^+]$.

Since acetic acid (CH_3COOH) is a monoprotic acid, $[H^+]$ = molar concentration of acid × degree of dissociation.

Therefore, $[H^+] = 1\times0.10\ \text{mol/L}\times0.016 = 1.6\times10^{-3}\ \text{mol/L}$

Therefore, $pH = -\log_{10}[H^+] = -\log_{10}(1.6\times10^{-3}) = -(\log_{10}1.6+\log_{10}10^{-3}) = -(0.2-3) = 2.8$

(b) (i) Let α be the degree of dissociation. Using the initial concentration C and the dissociation constant K, from the relationship

$$\alpha = \sqrt{\frac{K}{C}},$$

$$\alpha = \sqrt{\frac{1.8\times10^{-5}\ \text{mol/L}}{0.020\ \text{mol/L}}} = \sqrt{9.0\times10^{-4}} = \underline{3.0\times10^{-2}}$$

(ii) Since ammonia (NH_3) produces one hydroxide ion per molecule, $[OH^-]$ = molar concentration of base × degree of dissociation, and from (i), the degree of dissociation is 3.0×10^{-2}, so $[OH^-] = 1\times0.020\ \text{mol/L}\times3.0\times10^{-2} = \underline{6.0\times10^{-4}\ \text{mol/L}}$

(iii) Since pH = −log10[H+], first find the hydrogen ion concentration [H+].

From the ionic product of water $K_w = 1.0\times10^{-14}\ \text{mol}^2/\text{L}^2$, $[H^+][OH^-] = 1.0\times10^{-14}\ \text{mol}^2/\text{L}^2$

Therefore,

$$[H^+] = \frac{1.0\times10^{-14}\ \text{mol}^2/\text{L}^2}{[OH^-]} = \frac{1.0\times10^{-14}\ \text{mol}^2/\text{L}^2}{6.0\times10^{-4}\ \text{mol/L}} = \frac{1.0}{6.0}\times10^{-10}\ \text{mol/L}$$

Therefore, the required pH is $pH = -\log_{10}[H^+]$

$$= -\log_{10}\left(\frac{1.0}{6.0}\times10^{-10}\right)$$

$$= -\left\{\log_{10}\left(\frac{1.0}{6.0}\right)+\log_{10}10^{-10}\right\}$$

$$= -(\log_{10}6.0^{-1}+\log_{10}10^{-10})$$

$$= -(-\log_{10}6.0+\log_{10}10^{-10})$$

$$= \log_{10}6.0-\log_{10}10^{-10}$$

$$= \log_{10}(2.0\times3.0)-\log_{10}10^{-10}$$

$$= \log_{10}2.0+\log_{10}3.0-\log_{10}10^{-10}$$

$$= 0.30+0.48-(-10) = 10.78 \approx 10.8 \quad 10.8$$

Data Interpretation Activity

Substance	Initial Temp (°C)	Electrical Conductivity
HCl	High	Very High
CH_3COOH	High	Low
Glucose	High	None

• Think Like a Chemist:.

1. Which substance is a strong electrolyte? ....

2. Which is a weak electrolyte?..

3. Which is a non-electrolyte? ....

4. What evidence supports your classification? ....

Try

Choose the correct answer:

1. A strong electrolyte is best described as:

A. Partially ionized

B. Fully ionized in solution

C. Insoluble in water

D. Non-conducting

Answer: ___________________________________

2. Which substance is a weak electrolyte?

A. NaCl

B. HCl

C. CH_3COOH

D. Glucose

Answer: ___________________________________

Answer the following questions

3. Assume Kw = 1.0×10^{-14} mol^2/L^2, and calculate the pH to one decimal place. ($\log_{10} 2.0 = 0.3$, $\log_{10} 5.0 = 0.7$)

(a) Calculate the pH of 0.020 mol/L hydrochloric acid (degree of dissociation 1.0).

(b) Calculate the pH of a 0.050 mol/L aqueous acetic acid solution (degree of dissociation 0.050).

(c) Calculate the pH of a 0.025 mol/L aqueous barium hydroxide solution (degree of dissociation 1.0).

(d) Calculate the pH of 0.050 mol/L aqueous ammonia (degree of dissociation 0.020).

Answer: ___________________________________

Exercise

1. Calculate the pH of the following aqueous solutions at 25°C to one decimal place. Assume the ionic product of water is $K_w = 1.0\times10^{-14}$ mol^2/L^2. ($\log_{10}1.5 = 0.18$, $\log_{10}2.0 = 0.30$)

(a) 0.0050 mol/L sulfuric acid

Answer: ____

(b) 0.20 mol/L aqueous acetic acid solution (degree of dissociation 0.016)

Answer: ____

(c) 0.10 mol/L aqueous sodium hydroxide solution

Answer: ____

(d) 0.10 mol/L aqueous ammonia (degree of dissociation 1.5×10^{-2})

Answer: ____

2. In aqueous ammonia, the following dissociation equilibrium is established.

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

Assuming the dissociation constant of ammonia K_b is 1.8×10^{-5} mol/L, the ionic product of water K_w is 1.0×10^{-14} mol^2/L^2, and the degree of dissociation of ammonia is much smaller than 1, answer the following questions.

(a) Express the degree of dissociation α of ammonia in c [mol/L] aqueous ammonia using c and K_b.

Answer: ____

(b) Calculate the hydroxide ion concentration [OH^-] in 2.0 mol/L aqueous ammonia to two significant figures.

Answer: ____

(c) Calculate the pH of the aqueous ammonia in (b) to one decimal place. Assume $\log_{10}2 = 0.30$ and $\log_{10}3 = 0.48$.

Answer: ____

Challenge Your Thinking

A chemist compares two solutions:

• 0.1 M HCl

• 0.1 M CH_3COOH

Both are acids, but only one conducts strongly.

• Is conductivity determined by concentration alone? ....

• What role does ionization play?..

• Can weak acids become strong conductors? Under what condition? ....

• Explain using equilibrium reasoning. ....

Lesson 2-12 — Salt and Aqueous Solution

✨ Engaging Introduction

When you add hydrochloric acid to a solution of sodium acetate, at first glance nothing dramatic seems to happen. However, at the molecular level a hidden competition begins: ions collide, react, and shift equilibrium positions continuously.

In another case, adding sodium hydroxide to an ammonium salt produces a noticeable smell of ammonia gas — evidence that a weak base has been released from its salt form.

In this lesson, we will explore how equilibrium decides which acid or base survives in solution and how salts control the acidity or basicity of water.

❓ Guiding question: How does ionization equilibrium determine whether a weak acid or weak base is displaced in aqueous solution?

🔭 Exploration

Consider the following reactions:

- $CH_3COONa + HCl \rightarrow ?$
- $NH_4Cl + NaOH \rightarrow ?$

Without studying the lesson:

- What do you expect to form when a strong acid reacts with a salt of a weak acid?

....

- What do you expect to form when a strong base reacts with a salt of a weak base?

....

- Why do some salts produce acidic or basic solutions in water?

....

📌 Point!

1. Displacement of weak acids and weak bases

(a) Displacement of weak acids: When a strong acid is added to a salt of a weak acid, a salt of the strong acid and a weak acid are formed.

Example: $\underset{\text{Salt of weak acid}}{CH_3COONa} + \underset{\text{strong acid}}{HCl} \longrightarrow \underset{\text{Salt of strong acid}}{NaCl} + \underset{\text{Weak acid}}{CH_3COOH}$

(b) Displacement of weak bases: When a strong base is added to a salt of a weak base, a salt of the strong base and a weak base are formed.

Example: $\underset{\text{Salt of weak base}}{NH_4Cl} + \underset{\text{Strong base}}{NaOH} \longrightarrow \underset{\text{Salt of strong base}}{NaCl} + \underset{\text{Weak base}}{NH_3} + \underset{\text{Water}}{H_2O}$

⭐ **Structure of Science**

Key Fact :

Strong acids and bases can displace weak acids and weak bases from their salts. • Buffer solutions resist changes in pH due to equilibrium control of ions.

Key Concepts:

.Ionization equilibrium
· Acid–base reaction
· Salt hydrolysis
· Displacement reaction
· Weak acid
· Weak base
· Strong acid
· Strong base
· Buffer solution
· Conjugate acid–base pair
· pH stability

2. Salts and aqueous solutions

(a) Classification of salts

(i) Normal salt: A salt in which neither the H of the acid nor the OH of the base remains in the chemical formula. **Example:** Sodium carbonate Na_2CO_3

(ii) Acid salt: A salt in which the H of the acid remains. **Example:** Sodium hydrogen carbonate $NaHCO_3$

(iii) Basic salt: A salt in which the OH of the base remains. **Example:** Calcium chloride hydroxide CaCl(OH)

(b) Aqueous solutions of normal salts

Generally, the relationship between the composition of a normal salt and the property of its aqueous solution is as follows.

(i) Weak acid + Strong base $\longrightarrow$ Basic

(ii) Strong acid + Weak base $\longrightarrow$ Acid

(iii) Strong acid + Strong base $\longrightarrow$ Neutral

(c) Aqueous solutions of acid salts

(i) An aqueous solution of sodium hydrogen carbonate $NaHCO_3$ is basic .

This is because HCO_3^- generated by ionisation reacts with H_2O to produce OH^-.

(ii) An aqueous solution of sodium hydrogen sulfate $NaHSO_4$ is acidic .

This is because HSO_4^- generated by ionisation further ionises to produce H^+.

3. Buffer action

(a) Buffering action : The ability to resist changes in pH even when a small amount of acid or base is added.

(b) Buffer solution : An aqueous solution that exhibits buffer action. Generally, a mixed aqueous solution of a weak acid or a weak base and its salt is used.

(c) Mechanism of buffer action (in the case of a mixed aqueous solution of acetic acid and sodium acetate)

Acetic acid is a weak acid, and sodium acetate is a salt of acetic acid.

In the aqueous solution, acetic acid and sodium acetate are ionised.

$$\begin{cases} \text{Ionisation of acetic acid: } CH_3COOH \rightleftarrows CH_3COO^- + H^+ \\ \text{Ionisation of sodium acetate: } CH_3COONa \longrightarrow CH_3COO^- + Na^+ \end{cases}$$

Misconception Alert

❌ Wrong idea:

„All salts produce neutral solutions."

✅ Correct:

The pH of a salt solution depends on the strengths of the acid and base from which the salt is formed.

❌ Wrong idea:

„A buffer solution prevents acids and bases from reacting."

✅ Correct:

A buffer solution allows reactions to occur but minimizes changes in pH through equilibrium shifts.

(i) When H^+ is added, H^+ reacts with CH_3COO^- to form acetic acid molecules CH_3COOH.

$$CH_3COO^- + H^+ \longrightarrow CH_3COOH$$

⇒ $[H^+]$ hardly increases.

(ii) When OH^- is added, it reacts with acetic acid molecules to form water molecules H_2O.

$$CH_3COOH + OH^- \longrightarrow CH_3COO^- + H_2O$$

⇒ $[OH^-]$ hardly increases.

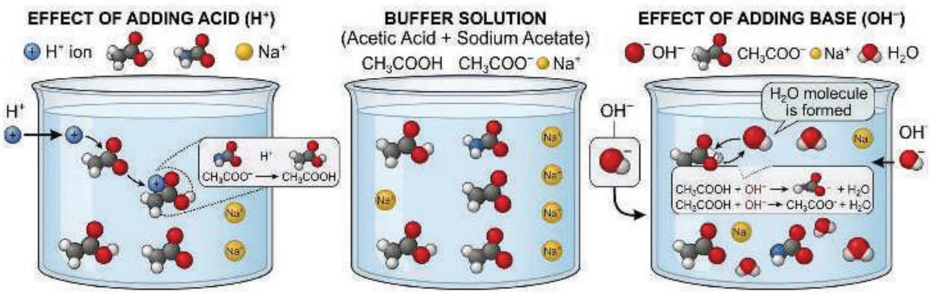


Figure 2.12.1 the three-panel buffer diagram

📦 Reference table — salts & buffers at a glance

Normal salt from…	Solution property
Strong acid + Strong base	Neutral
Strong acid + Weak base	Acidic
Weak acid + Strong base	Basic
Acid salt — $NaHCO_3$ / $NaHSO_4$	basic / acidic
Buffer (weak acid + its salt)	resists pH change
Add H^+ to buffer	$CH_3COO^- + H^+ \rightarrow CH_3COOH$
Add OH^- to buffer	$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

Warm Up

Answer the following questions.

(a) Write the chemical equation for the reaction that occurs when hydrochloric acid is added to an aqueous sodium hydrogen carbonate solution.

(b) State whether the aqueous solutions of the following salts are acidic, basic, or neutral.

(i) NaCl (ii) $Ba(NO_3)_2$ (iii) NH_4Cl (iv) CH_3COONa

(c) Insert appropriate ionic equations into [A] and [B], and an appropriate word into (C) in the following text. Also, choose appropriate items for { D } and { E } from (i) to (v) below.

When a small amount of hydrochloric acid is added to a mixed aqueous solution containing equimolar amounts of acetic acid and sodium acetate, the reaction of

🌍 Be Global Citizen

Understanding buffer systems helps scientists:

- Develop safer medicines and intravenous solutions.
- Monitor and protect aquatic ecosystems from acidification.
- Improve food preservation technologies.
- Maintain stable conditions in industrial and biological processes.

Many sustainability and healthcare applications depend on controlling ionization equilibrium and buff.

[A] occurs, and the hydrogen ion concentration is kept almost constant. Also, when a small amount of aqueous sodium hydroxide solution is added, the reaction of [B] occurs, and the hydroxide ion concentration is kept almost constant. Such an aqueous solution is called a (C), and a similar phenomenon occurs in a mixed aqueous solution containing equimolar amounts of the substances in { D } and { E }.

(i) NH_3 (ii) HCl (iii) NaCl (iv) NH_4Cl (v) CH_3COOH

Explanation

(a) Since sodium hydrogen carbonate is a salt of a weak acid and hydrochloric acid is a strong acid, sodium chloride, which is a salt of a strong acid, is formed, and the weak acid, carbonic acid, is displaced.

$$NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$$

(b) (i) NaCl is a normal salt derived from the strong acid HCl and the strong base NaOH. Therefore, it is neutral .

(ii) $Ba(NO_3)_2$ is a normal salt derived from the strong acid HNO_3 and the strong base $Ba(OH)_2$. Therefore, it is neutral .

(iii) NH_4Cl is a normal salt derived from the strong acid HCl and the weak base NH_3. Therefore, it is acidic .

(iv) CH_3COONa is a normal salt derived from the weak acid CH_3COOH and the strong base NaOH.
Therefore, it is basic .

(c) A: $CH_3COO^- + H^+ \longrightarrow CH_3COOH$

B: $CH_3COOH + OH^- \longrightarrow CH_3COO^- + H_2O$

C: Buffer solution

D, E: A mixed aqueous solution of ammonia NH_3, which is a weak base, and ammonium chloride, which is a salt of ammonia, exhibits buffering action. (i), (iv) (in any order)

Data Interpretation Activity

Solution	Initial pH	Final pH after adding HCl
Distilled Water	7.0	3.0
CH_3COOH / CH_3COONa Buffer	4.8	4.7

Analyze Experimental Data

Think Like a Chemist

1. Which solution resisted the pH change most effectively? ....

2. What evidence supports your conclusion? ....

3. How did the equilibrium respond to the added acid? ....

4. Which solution demonstrates buffer action? ....

Try

1. Write the balanced chemical equation for each of the following reactions.

(a) Add hydrochloric acid to an aqueous potassium acetate solution.
Answer: ______________________________________

(b) Add an aqueous calcium hydroxide solution to an aqueous ammonium chloride solution.
Answer: ______________________________________

2. State whether the aqueous solutions of the following salts are acidic, basic, or neutral.

(a) KNO_3 (b) $(NH_4)_2SO_4$ (c) Na_2CO_3 (d) $NaHSO_4$

3. Read the following text and answer the questions.

Acetic acid is a weak acid, slightly ionized in an aqueous solution, maintaining the following equilibrium.

$CH_3COOH \rightleftarrows CH_3COO^- + H^+$ ......(i)

When sodium acetate is added to this aqueous solution, the salt completely dissociates, so the concentration of [A] increases, and according to Le Chatelier's principle, the equilibrium shifts to the (B). As a result, the concentration of H^+ decreases, resulting in an aqueous solution in which large amounts of [C] and [D] are present.

When a small amount of hydrochloric acid is added to this mixed aqueous solution, the chemical equilibrium of equation (i) shifts to the (E) side, and (I)[F] generated from hydrochloric acid combines with the large amount of CH_3COO^- present to become [G]. Therefore, the concentration of [H] in the mixed aqueous solution hardly changes.

Also, when a small amount of sodium hydroxide is added to the mixed aqueous solution, (II)[I] generated from sodium hydroxide reacts with CH_3COOH. Therefore, the concentration of [J] in the aqueous solution hardly changes. This function of keeping the pH of the solution almost constant is called (K).

(a) Insert appropriate chemical formulas into [ ], and appropriate words into ( ) in the text. The same item may be inserted multiple times.
Answer: ______________________________________

(b) Write the net ionic equation for the reaction in the underlined part I.
Answer: ______________________________________

(c) Write the net ionic equation for the reaction in the underlined part II.
Answer: ______________________________________

Think Critically

A scientist compares two solutions:

- Solution A: Acetic acid only
- Solution B: Acetic acid + sodium acetate

The same amount of hydrochloric acid is added to both solutions.

Evaluate the situation

1. Which solution will experience the smaller pH change? ....
2. Which ion is primarily responsible for resisting the increase in H^+ concentration?

....

3. How does equilibrium shift after adding H^+ ions? ....
4. Why is the buffer solution considered a dynamic equilibrium system? ....

Exercise

Choose the correct answer:

1. Which salt produces an acidic aqueous solution?

A. $NaCl$

B. $Ba(NO_3)_2$

C. NH_4Cl

D. CH_3COONa

Answer: ____

2. Which ionic equation explains the action of an acetate buffer when acid is added?

A. $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

B. $CH_3COO^- + H^+ \rightarrow CH_3COOH$

C. $H^+ + OH^- \rightarrow H_2O$

D. $NH_4^+ + OH^- \rightarrow NH_3 + H_2O$

Answer: ____

Answer on the following questions

3. Write the balanced chemical equation for the reaction between the following substances.

(a) CH_3COONa, H_2SO_4

(b) NH_4Cl, $NaOH$

Answer: ____

4. Read the following text and answer the questions. (I) When sodium acetate is dissolved in water, it dissociates into acetate ions and sodium ions. A part of the acetate ions reacts with water molecules to produce hydroxide ions, so the aqueous solution is weakly (A).

(a) Insert an appropriate word into (A). Also write the net ionic equation for the reaction in the underlined part (I).

Answer: ____

(b) State whether each of the following aqueous solutions is acidic, basic, or neutral.

Answer: ____

(i) Equal volumes of 0.10 mol/L hydrochloric acid and 0.10 mol/L potassium hydroxide.

Answer: ____

(ii) Equal volumes of 0.10 mol/L acetic acid and 0.10 mol/L potassium hydroxide.

Answer: ____

(iii) Equal volumes of 0.10 mol/L hydrochloric acid and 0.10 mol/L aqueous ammonia.

Answer: ____

Lesson 2-13 — Solubility Product

✨ Engaging Introduction

Two solutions—one clear and colorless, and the other yellow are mixed together in a laboratory. At first, nothing seems unusual. Suddenly, a brick-red solid appears and settles at the bottom of the beaker. Where did this solid come from?

At the microscopic level, dissolved ions are constantly moving through the solution. Sometimes they remain dissolved, while at other times they combine to form a precipitate. Chemists use the concept of solubility equilibrium and the solubility product constant (Ksp) to predict exactly when precipitation will occur.

In this lesson, we will explore how equilibrium controls the balance between dissolved ions and solid substances and how scientists use Ksp to predict chemical changes in natural and industrial systems.

> ❓ **Guiding question:** How can chemists predict whether dissolved ions will remain in solution or form a precipitate?

🔭 Predict what happens

A solution containing Ag^{+} ions is mixed with a solution containing CrO_4^{2-} ions.

Without studying the lesson:

• Will the ions remain dissolved?

……

• Could a solid form even though both solutions were initially clear?

……

• What factors might determine whether precipitation occurs?

……

Record your ideas and revisit them after studying solubility equilibrium..

📌 **Point!**

Solubility product

(a) Solubility equilibrium : A state in a saturated solution where the rate of dissolution equals the rate of precipitation, and macroscopically no net change is observed.

(b) Solubility product : The product of the concentrations of the dissolved ions at solubility equilibrium. It is denoted by the symbol K_{sp} and remains constant as long as the temperature does not change.

(c) Expression for Solubility Product

When a sparingly soluble salt A_aB_b consisting of A^{m+} and B^{n-} is in the following equilibrium,

⭐ **Structure of Science**

Key Fact :

Solubility equilibrium exists when the rate of dissolution equals the rate of precipitation.

• The solubility product constant (Ksp) depends only on temperature.

Key Concepts:

Solubility equilibrium · Solubility product (Ksp)

· Saturated solution

· Dynamic equilibrium

· Dissolution

· Precipitation

· Ion product

· Ionic concentration · Sparingly soluble salt.

$A_aB_b \rightleftarrows aA^{m+} + bB^{n-}$ (where $am = bn$)

the solubility product K_{sp} is expressed as follows.

$K_{sp} = [A^{m+}]^a[B^{n-}]^b$

Example: For silver chromate Ag_2CrO_4 , when the solubility equilibrium

$Ag_2CrO_{4(s)} \rightleftarrows 2Ag^+ + CrO_4^{2-}$ is established,

Solubility product K_{sp} = $[Ag^+]^2[CrO_4^{2-}]$

(d) Solubility Product and Precipitation

Whether precipitation occurs in an aqueous solution is considered by comparing the product of the molar concentrations of ions with the solubility product.

Example: In the case of silver chromate Ag_2CrO_4

When $[Ag^+]^2[CrO_4^{2-}]$ $\leq$ K_{sp}, precipitation does not occur.

When the equality holds, a dissolution equilibrium is established.

When $[Ag^+]^2[CrO_4^{2-}]$ $>$ K_{sp}, precipitation occurs.

Warm Up

Hydrogen sulfide H_2S was passed through an aqueous solution containing 1.0×10^{-4} mol/L of lead(II) ions Pb^{2+} resulting in a sulfide ion concentration of 1.0×10^{-20} mol/L. Answer the following questions.

(a) Write the ionic equation at solubility equilibrium.

(b) Express the solubility product K_{sp}.

(c) Does precipitation of lead(II) sulfide PbS occur? Assume the solubility product of PbS is $K_{sp} = 1.0\times10^{-28}$ mol²/L².

Explanation

(a) Hydrogen sulfide ionises to produce H^+ and S^{2-}. In the aqueous solution, Pb^{2+} and S^{2-} react to form PbS.

$PbS(s) \rightleftarrows Pb^{2+} + S^{2-}$

(b) From **(a)**, since all the coefficients of the chemical equation are 1, $K_{sp} = [Pb^{2+}][S^{2-}]$

(c) Compare the product of the molar concentrations of ions, which is expressed by the same formula as K_{sp} , with K_{sp} .

Assuming precipitation of PbS does not occur, from

$[Pb^{2+}] = 1.0\times10^{-4}$ mol/L and $[S^{2-}] = 1.0\times10^{-20}$ mol/L,

$[Pb^{2+}][S^{2-}] = 1.0\times10^{-4}\ \text{mol/L}\times1.0\times10^{-20}\ \text{mol/L} = 1.0\times10^{-24}\ \text{mol}^2/\text{L}^2$

Since $K_{sp} = 1.0\times10^{-28}$ mol²/L² , $[Pb^{2+}][S^{2-}]>K_{sp}$

Therefore, precipitation occurs.

Misconception Alert

❌ Wrong idea:

"If a solid is visible in a solution, equilibrium has stopped."

✅ Correct:

Equilibrium continues because dissolution and precipitation occur simultaneously at equal rates.

Data Interpretation Activity

Analyze Experimental Data

Solution	$[Ag^+]$ (mol/L)	$[CrO_4^{2-}]$ (mol/L)	Ion Product
A	1.0×10^{-4}	1.0×10^{-4}	1.0×10^{-12}
B	1.0×10^{-3}	1.0×10^{-3}	1.0×10^{-9}

Assume:

Ksp (Ag_2CrO_4) = 1.0×10^{-11}

Think Like a Chemist

1. In which solution is precipitation more likely? ....
2. Which solution has an ion product greater than Ksp? ....
3. What evidence supports your conclusion? ....
4. What happens when the ion product exceeds Ksp? ....

Try

In an aqueous solution where the molar concentration of magnesium ions (Mg^{2+}) is 1.0×10^{-3} mol/L, precipitation of magnesium hydroxide $Mg(OH)_2$ begins to form when the molar concentration of hydroxide ions (OH^-) exceeds x [mol/L]. Assume the solubility product of $Mg(OH)_2$ is 1.0×10^{-11} $[mol/L]^3$. Answer the following questions.

(a) Write the ionic equation for the reaction when precipitation forms and solubility equilibrium is reached.

(b) Write the solubility product expression for K_{sp}.

(c) Determine the value of x to two significant figures.

Answer: ______________________________________

Think Critically

A factory releases wastewater containing Pb^{2+} ions into a treatment system.

A scientist proposes adding sulfide ions (S^{2-}) to remove lead ions as solid PbS.

Evaluate the proposal.

1. Why might PbS precipitation occur? ....
2. What role does Ksp play in this process? ....
3. How can precipitation reduce environmental pollution? ....
4. What evidence would confirm that Pb^{2+} ions have been removed? ....

Exercise

The solubility of silver chloride at 25°C is 1.9 mg/L. Answer the following questions to two significant figures. (Na = 23, Cl = 35.5, Ag = 108)

(a) What is the molar concentration of Ag^{+} in a saturated solution of silver chloride at 25°C [in mol/L]?

(b) Determine the solubility product of silver chloride at 25°C.

(c) When sodium chloride is added gradually to 100 mL of a 1.0×10^{-3} mol/L aqueous silver nitrate solution, what is the minimum mass in milligrams of sodium chloride that must be added to initiate the precipitation of silver chloride?

Answer: ___________________________________

Unit question

Sustainable Energy Systems: Choosing the Best Technology for a Green City

A city plans to reduce its carbon emissions by replacing conventional energy systems with cleaner technologies. Engineers are evaluating three technologies:

1. Hydrogen Fuel Cells

2. Lead-Acid Batteries

3. Hydrogen Production by Electrolysis

The following data have been collected.

Data Set 1: Energy and Environmental Information

Technology	Main Chemical Process	ΔH per mole of reaction".	Main Product
Hydrogen Fuel Cell	$2H_2 + O_2 \rightarrow 2H_2O$	−571.6 kJ	Water
Lead-Acid Battery (Discharge)	$Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$	Releases energy	Water and $PbSO_4$
Water Electrolysis	$2H_2O \rightarrow 2H_2 + O_2$	+572	Hydrogen and Oxygen

Data Set 2: Electrical Charge Production

System	**Moles of Electrons Transferred**
System A	1.0 mol
System B	2.0 mol
System C	5.0 mol

....

....

Section A: Multiple-Choice Question

Which statement is best supported by the data?

A. Water electrolysis is exothermic because hydrogen is produced.

B. Fuel cells and electrolysis have identical energy changes because they are the same process.

C. Water electrolysis requires energy input, while fuel cells release energy.

D. Lead-acid batteries and fuel cells both consume water as a reactant.

Answer

....

Justification

....

Section B: Short Response

Define enthalpy change (ΔH) and explain the difference between an exothermic and an endothermic process using the data provided.

....

Section C: Structured Problem

Calculate and Explain

Using Data Set 2:

(a) Calculate the electrical charge produced by each system.

(b)Identify which system could produce the greatest amount of chemical change.

(c)Explain your reasoning using Faraday's Law.

....

Section D: Data-Response Analysis

A student makes the following claim:

"Hydrogen fuel cells should completely replace lead-acid batteries because they produce only water."

Evaluate this claim.

In your answer:

- discuss scientific advantages,
- discuss limitations,
- consider energy production, storage, and sustainability,
- use evidence from the case study.

....

Section E: Extended Essay

To what extent can electrochemical technologies contribute to a sustainable energy future?

In your response:

- Explain the role of redox reactions in electrochemical cells.

....

- Compare fuel cells, lead-acid batteries, and electrolysis.

....

- Analyze the relationship between energy changes (ΔH) and electrical energy production.

....

- Discuss the environmental and economic implications of each technology.

....

- Use scientific evidence to support your conclusions.

....

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