Surfing

UNIT 3

QCE CHEMISTRY

Unit 3 Equilibrium, Acids and Redox Reactions

• Marilyn Schell • Margaret Hogan •

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Unit 7, 23-31 Bowden Street
Alexandria NSW 2015 Australia
Tel: +61 2 9020 1840 Fax: +61 2 9020 1842
sales@sciencepress.com.au
www.sciencepress.com.au

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Introduction

This book covers the Chemistry content specified in the Queensland Certificate of Education Chemistry Syllabus. Sample data has been included for suggested experiments to give you practice to reinforce practical work in class.

Each book in the *Surfing* series contains a summary, with occasional more detailed sections, of all the mandatory parts of the syllabus, along with questions and answers.

All types of questions – multiple choice, short response, structured response and free response – are provided. Questions are written in exam style so that you will become familiar with the concepts of the topic and answering questions in the required way.

Answers to all questions are included.

A topic test at the end of each topic contains an extensive set of summary questions. These cover every aspect of the topic, and are useful for revision and exam practice.

Words To Watch

account, account for State reasons for, report on, give an account of, narrate a series of events or transactions.

analyse Interpret data to reach conclusions.

annotate Add brief notes to a diagram or graph.

apply Put to use in a particular situation.

assess Make a judgement about the value of something.

calculate Find a numerical answer.

clarify Make clear or plain.

classify Arrange into classes, groups or categories.

comment Give a judgement based on a given statement or result of a calculation.

compare Estimate, measure or note how things are similar or different.

construct Represent or develop in graphical form.

contrast Show how things are different or opposite.

create Originate or bring into existence.

deduce Reach a conclusion from given information.

define Give the precise meaning of a word, phrase or physical quantity.

demonstrate Show by example.

derive Manipulate a mathematical relationship(s) to give a new equation or relationship.

describe Give a detailed account.

design Produce a plan, simulation or model.

determine Find the only possible answer.

discuss Talk or write about a topic, taking into account different issues or ideas.

distinguish Give differences between two or more different items.

draw Represent by means of pencil lines.

estimate Find an approximate value for an unknown quantity.

evaluate Assess the implications and limitations.

examine Inquire into.

explain Make something clear or easy to understand.

extract Choose relevant and/or appropriate details.

extrapolate Infer from what is known.

hypothesise Suggest an explanation for a group of facts or phenomena.

identify Recognise and name.

interpret Draw meaning from.

investigate Plan, inquire into and draw conclusions about.

justify Support an argument or conclusion.

label Add labels to a diagram.

list Give a sequence of names or other brief answers.

measure Find a value for a quantity.

outline Give a brief account or summary.

plan Use strategies to develop a series of steps or processes.

predict Give an expected result.

propose Put forward a plan or suggestion for consideration or action.

recall Present remembered ideas, facts or experiences.

relate Tell or report about happenings, events or circumstances.

represent Use words, images or symbols to convey meaning.

select Choose in preference to another or others.

sequence Arrange in order.

show Give the steps in a calculation or derivation.

sketch Make a quick, rough drawing of something.

solve Work out the answer to a problem.

state Give a specific name, value or other brief answer.

suggest Put forward an idea for consideration.

summarise Give a brief statement of the main points.

synthesise Combine various elements to make a whole.

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QCE CHEMISTRY

UNIT 3

UNIT 3

Equilibrium, Acids and Redox Reactions

In this unit you will:

- Explore the reversibility of chemical reactions.
- Explore acid-base equilibrium systems, oxidation and reduction, and the production of electricity from electrochemical cells.
- Use contemporary models to explain the nature of acids and bases, their properties and uses.
- Investigate volumetric analysis and its applications.
- Use contexts to investigate environmental issues and the historical development of theories.
- Explore the contributions of chemistry in industrial and environmental contexts.
- Understand the pH scale is a logarithmic scale.
- Solve problems for $[H^+]$, $[OH^-]$, pH and pOH.

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QCE CHEMISTRY

UNIT 3

TOPIC 7

Chemical Equilibrium Systems

In this topic you will:

- Understand the concepts of open and closed systems, dynamic equilibrium, and activation energies.
- Explain the effects of changes in temperature, concentration, pressure and catalysts on an equilibrium system.
- Apply Le Châtelier's principle to predict the effects of changes on equilibrium.
- Use the equilibrium constant (K_c) to predict equilibrium position and the extent of a reaction and solve problems.
- Understand that acids act as proton donors.
- Predict, explain and distinguish between strong and weak acids.
- Use the self-ionisation of water K_w to calculate concentrations of H^+ and OH^- ions.
- Use the Brönsted-Lowry model to explain acids and bases, amphiprotic species and conjugate acids/ bases.
- Apply Le Châtelier's principle to predict how buffer solutions respond to additions of H^+ and OH^- ions.
- Determine dissociation constants K_a and K_b.
- Explain the relationship between pH range of an indicator and pK_a value.
- Carry out and analyse acid-base titrations as a volumetric analysis technique using appropriate indicators.
- Use the mole concept and volumetric analysis data in calculations.

1 Open and Closed Systems

In this course you will be looking at chemical equilibrium systems. So we will look first at what is meant by a system, and then what is meant by equilibrium.

You can think of a **system** as any part of the Universe that is being studied. Sometimes a system can be very large, such as a whole ocean. A system can also be relatively small, such as the contents of a test tube.

In chemistry, the systems we study can be either open or closed; they can involve either physical changes or chemical reactions; and the changes taking place can either go to completion or be reversible.

Open and closed systems

An **open system** is one which interacts with its environment – this means that both energy and matter can move in and out of the system. Energy and matter are constantly moving between the system and the environment.

In a **closed system**, energy may still be able to flow in and out, between the system and the environment, but matter cannot enter or leave the system. When a reaction reaches equilibrium, no energy enters or leaves the system either. (If you are also studying physics, note that this definition is different in physics.)

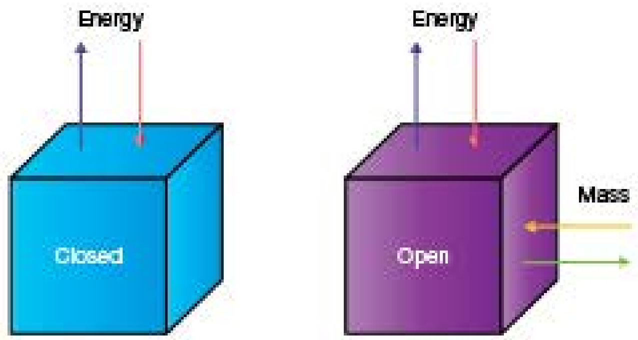


Figure 1.1 Open and closed systems.

A physical change

You will recall that a physical change is one in which no new substances are made, for example a change of state. The particles stay the same, but they may move differently, speeding up or slowing down. This may occur in an open or a closed system.

If we heat water in a container without a lid, then we have a **physical change** (evaporation) taking place in an **open system**.

If we put the lid on, then we have the same **physical change** taking place, but in a **closed system**. Heat can go in, but no matter can enter or leave the container.

Figure 1.2 A physical change in an open and a closed system.

A chemical change

You will recall that a chemical change is one in which one or more new substances are formed, for example combustion, the action of acids on active metals, synthesis and decomposition.

Chemical changes can also occur in both open and closed systems.

Reversible or irreversible?

Physical changes are usually reversible, they can proceed in both directions. For example, ice (solid water) will melt when warmed, and we show this as:

$H_2O(s) \rightarrow H_2O(l)$

If we cool the liquid water it can freeze and once again become solid:

$H_2O(l) \rightarrow H_2O(s)$

This is a reversible physical change. No new substance is produced, it is always water. It is just changing state.

To show this as a reversible reaction we use a two-way arrow to show that the reaction can proceed in either direction.

$H_2O(s) \rightleftharpoons H_2O(l)$

Chemical reactions are usually not reversible under normal circumstances. Most reactions that you have seen will proceed in one direction only, they go to completion and they cannot be reversed. We describe these reactions as irreversible.

You have seen many irreversible reactions in the laboratory, for example, neutralisation reactions between metals and acids and combustion reactions. Cooking a cake involves irreversible chemical reactions. You cannot turn a cake back into its original ingredients – such as flour, eggs, sugar and milk.

Reversible chemical reactions

There are many chemical reactions that are reversible in a closed system and we will be investigating such reversible chemical reactions in this unit.

In a reversible chemical reaction, the reaction can go in both directions, and both the forward and reverse reactions can take place at the same time in the same container.

When you mix reactants, under suitable conditions, as the products build up, they may start to re-form the reactants.

Both the forward and reverse reactions will then take place at the same time. This can only happen if nothing escapes from the container. If for example a gas escapes, both reactions cannot occur. Remember, the reaction system must be closed to be reversible.

As in physical changes, we show a reversible chemical reaction by using arrows going in both directions. Here are some examples of reversible chemical reactions.

- **Dinitrogen tetroxide** is a colourless gas that decomposes at room temperature to form brown nitrogen dioxide gas. In a sealed container, the reverse reaction also occurs, with some nitrogen dioxide combining to re-form dinitrogen tetroxide.

  $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

- An important reversible reaction that you learned about in year 11 when you studied acids, is the **ionisation of water** to form hydrogen ions and hydroxide ions.

  $H_2O(aq) \rightleftharpoons H^+(aq) + OH^-(aq)$

- Many reversible reactions occur in **living organisms**, for example, reactions involved in the transport of oxygen and carbon dioxide around the body in blood.

  Haemoglobin + $O_2(g) \rightleftharpoons$ oxyhaemoglobin

  $CO_2(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$

- We also make use of some reversible reactions. For example, some eyeglasses have **photochromatic (transition) lenses** that darken in ultraviolet light. These work by means of a reversible reaction. They contain substances such as a silver halide which darkens in UV light, making the lens darker. Indoors the reaction reverses and the lenses become clear again.

Some reactions that go to completion can, under the right conditions, become reversible. For example, at some stage you will have heated calcium carbonate and watched it decompose into calcium oxide and carbon dioxide.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

You probably bubbled the gas produced through limewater, which turned milky, showing that the gas was carbon dioxide.

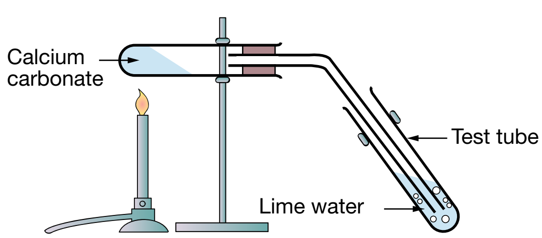


Figure 1.3 Decomposing calcium carbonate.

Note that this is an open system. There is a stopper in the test tube, but the carbon dioxide gas produced is escaping from the system through the tubing.

Heat energy is going into the system and matter (CO_2 gas) can escape. This is a **chemical change** in an **open system**.

However, if you were able to heat calcium carbonate gently in a closed container, heat energy would be going in, but no reactants or products could enter or leave. This would be a **chemical change** in a **closed system**.

This is not something you would do in a school laboratory. You would never enclose a reaction that produces a gas as this would not be safe. However, in an industrial situation, with containers that can withstand pressure, this reaction could be enclosed. In that case, some of the products – calcium oxide and carbon dioxide – would recombine to form calcium carbonate. The reverse reaction would occur.

$$CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$$

Both forward and reverse reactions would be taking place at the same time, and we would have a reversible reaction in a closed system.

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

QUESTIONS

1. Recall the differences between a physical and a chemical change.
2. Describe the meaning of the following terms.
   (a) A system.
   (b) Open system.
   (c) Closed system.
   (d) Reversible reaction.
3. Some chemical reactions are reversible.
   (a) Write word equations for the following reversible reactions.
      (i) $FeCl_3(s) + 3H_2O(l) \rightleftharpoons Fe(OH)_3(s) + 3HCl(aq)$
      (ii) $CuSO_4.5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(l)$
      (iii) $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$
      (iv) $CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$
      (v) $3O_2(g) \rightleftharpoons 2O_3(g)$
   (b) Write symbolic equations for the following reversible reactions.
      (i) Sulfur dioxide + oxygen $\rightleftharpoons$ sulfur trioxide
      (ii) Nitrogen + hydrogen $\rightleftharpoons$ ammonia
      (iii) Nitrogen dioxide $\rightleftharpoons$ dinitrogen tetroxide
      (iv) Phosphorus pentachloride $\rightleftharpoons$ phosphorus trichloride + chlorine
      (v) Hydrogen + iodine $\rightleftharpoons$ hydrogen iodide

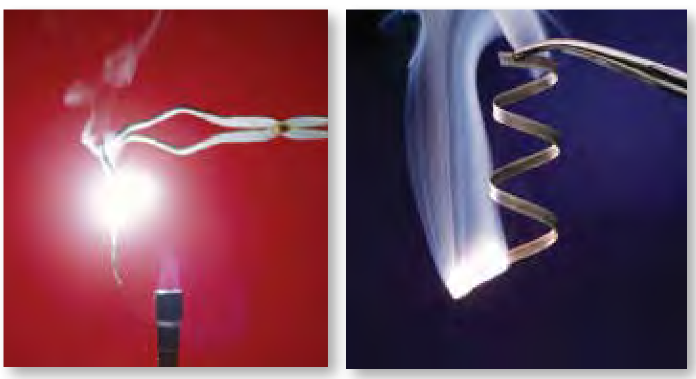
2 Investigating the Reversibility Of Chemical Reactions

In this topic you will have a chance to carry out reactions that are reversible – reactions that can proceed in either direction depending on prevailing conditions. You will learn to predict the direction in which a reaction will proceed by evaluating the conditions such as temperature, concentration and pressure to see if they favour the formation of reactants or products.

Irreversible reactions

In the laboratory you have carried out many reactions that proceed to completion – they are irreversible. For example, all combustion reactions are irreversible, they cannot be reversed.

One of the irreversible **combustion** reactions you have already seen is the combustion of **magnesium** ribbon. You know that energy is needed to start this reaction, and once started it continues with no more energy input as this is a highly exothermic reaction.



Bunsen starts magnesium burning

Magnesium continues to burn, giving out heat and a brilliant white light

Figure 2.1 Burning magnesium.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

This reaction cannot be reversed. You cannot convert the magnesium oxide produced back to the original elements magnesium and oxygen by changing conditions such as temperature – it is an irreversible reaction.

Steel wool is another substance that can be burnt in the laboratory in a non-reversible reaction. In the presence of oxygen, the thin pieces of steel burn to produce iron oxide. In pure oxygen, the reaction is much faster and more vigorous.

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Figure 2.2 Burning steel wool in oxygen.

Reversible reactions

Many transition metals form coloured salts, and their colour can change as they become hydrated or dehydrated. These are reversible reactions.

Copper chloride and iron chloride are examples as you can see in Figure 2.3.

(a) (b)

Figure 2.3 (a) Copper(I) chloride (CuCl) is green and copper(II) chloride is blue. (b) Crystals of iron(II) chloride tetrahydrate ($FeCl_2.4H_2O$) are green and crystals of iron(III) chloride hexahydrate ($FeCl_3.6H_2O$) are mustard yellow.

Cobalt chloride is another transition metal salt that changes colour. When hydrated its crystals are pink/purple ($CoCl_2.6H_2O$), when dehydrated it becomes anhydrous (without water) and is blue in colour. Cobalt chloride can be dried by heating the crystals to evaporate the water or by placing it in a desiccator. (If you are unfamiliar with a desiccator, see Question 3.)

(a) (b)

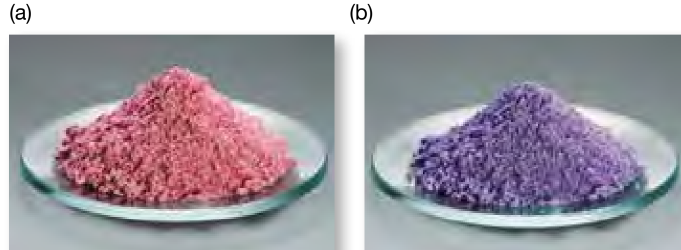


Figure 2.4 (a) Hydrated cobalt(II) chloride is pink/purple. (b) Anhydrous cobalt(II) chloride ($CoCl_2$) is blue.

Cobalt chloride paper

Strips of paper can be soaked in cobalt chloride solution to form cobalt chloride paper. This is blue when dried and it can be stored in a dry place such as a desiccator. It can be used to test for water leaks from tiny cracks – if the paper is applied and goes pink, then water is present.

Figure 2.5 Cobalt chloride paper goes pink in the presence of water.

Simple weather indicators can be made by soaking blotting paper in cobalt chloride and cutting it into interesting shapes such as the petals of a flower. It changes colour as the atmospheric humidity changes.

Another reversible reaction you will carry out involves **iron(III) nitrate and potassium thiocyanate**. When solutions of these ionic compounds are mixed, they react to form a blood-red complex called iron(III) thiocyanate. This is sometimes used as fake blood.

The nitrate ions and potassium ions are spectator ions and do not take part in the reaction, so the reaction is:

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons [FeSCN]^{2+}$$

Pale yellow colourless brown/red

This is a reversible reaction as both the forward and reverse reactions are occurring at the same time. The reactants do not all get used up. At any instant all species (reactants and products) are present in the reaction vessel.

Adding extra iron ions or thiocyanate ions separately, makes the colour a deeper red, as they react with ions that have not been used up.

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightarrow [FeSCN]^{2+}$$

Pale yellow colourless brown/red

Adding extra iron thiocyanate to the mixture produces a lighter colour as the reaction is pushed in the reverse direction.

$$[FeSCN]^{2+} \rightarrow Fe^{3+}(aq) + SCN^{-}(aq)$$

Brown/red pale yellow colourless

These changes could only happen if the reaction had not gone to conclusion – the reaction vessel must contain some reactants and also some product as both the forward and reverse reactions are happening simultaneously.

QUESTIONS

1. Describe the safety precautions you used when carrying out experiments involving:
   (a) The combustion of substances such as magnesium and steel wool.
   (b) The use of chemicals such as salts of transition elements.
2. (a) Describe how you made cobalt chloride change colour in the laboratory.
   (b) What happens to blue copper sulfate crystals when they are heated in a crucible above a Bunsen burner. Can this change be reversed?
   (c) Research uses of cobalt chloride paper.
3. The diagram illustrates a desiccator found in most school laboratories. What is a desiccator and what is it used for?

4. (a) In the reaction between iron nitrate and potassium thiocyanate, how could you determine which ions are spectator ions?
   (b) Research why iron thiocyanate can have a colour similar to blood.
5. Some students made up aqueous solutions of Fe^{3+} and SCN^{-} ions. They added these solutions and formed deep red iron(III) thiocyanate.
   Explain the following observations as they continued their investigation.
   (a) They added more of the solution containing Fe^{3+}, and the colour went a deeper red.
   (b) They also added more SCN^{-} and again the colour went a deeper red.
   (c) Extra $[FeSCN]^{2+}$ was added, and the mixture became lighter in colour.

3 Characteristics Of Equilibrium

You have seen that some reactions can proceed in both a forward and reverse direction at the same time. If such a reaction occurs in a **closed system** and the **forward and reverse reactions occur at the same rate**, then the reaction is said to be at **equilibrium**.

For different reactions, reaching equilibrium may take anything from milliseconds to many years.

If you look at a system in equilibrium, it will seem that nothing is happening. You will not see any changes taking place. There will be no change in the state (solid, liquid or gas) of the reactants or products. Also the pH, colour, temperature and pressure of the reactants and products will stay the same. Properties such as these that we can see are called macroscopic properties, so we say that at equilibrium the **macroscopic properties are constant**. This situation will continue unchanged as long as the external conditions, such as temperature, do not change and the system remains closed.

However, if you could see the tiny particles – the atoms and molecules – inside the container, you would see that there is constant change happening all the time. Reactions are continuously taking place at a microscopic level. We do not see these changes with our eyes because both the forward and reverse reactions are occurring at the same rate, so there is **no net change**.

Homogeneous and heterogeneous equilibrium

The term **homogeneous equilibrium** refers to a state of equilibrium in which all the involved species (reactants and products) are in the same state/phase. For example, the ionisation of a weak acid such as acetic acid is a homogeneous equilibrium reaction.

$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$

The Haber process used to make ammonia is another example of a homogeneous equilibrium reaction.

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

In a **heterogeneous equilibrium** reaction, substances can be in different states. For example, the decomposition of sodium hydrogen carbonate.

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

Figure 3.1 shows an example of a saturated solution of nickel chloride in a state of heterogeneous equilibrium. Crystals are forming on the bottom of the container and also going into solution, but both reactions are occurring at the same rate, so the amount of solid nickel chloride deposited is unchanged. As each nickel chloride crystal forms another one decomposes.

$NiCl_2(s) \rightleftharpoons Ni^{2+}(aq) + 2Cl^-(aq)$

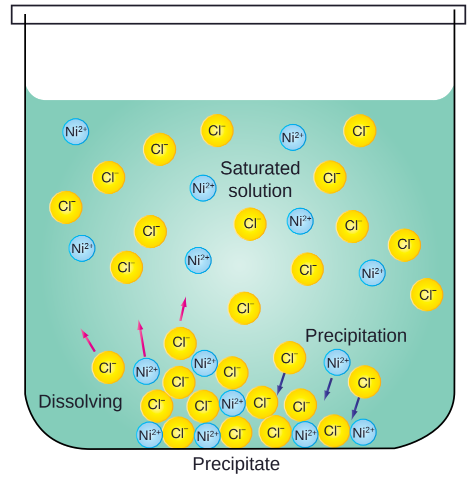


Figure 3.1 Saturated solution of nickel(II) chloride solution in equilibrium.

Characteristics of equilibrium systems

The **characteristics of an equilibrium system** are:

1. It is a **closed system**. No matter enters or leaves the system. When a closed system is at equilibrium, there can be no overall energy changes either. If a **reversible reaction** is to reach and maintain equilibrium, then **no matter or energy can enter or leave** the system. If the energy changes in a closed system, the equilibrium will change. So there cannot be equilibrium if energy is lost or gained from the system.
2. It involves a reversible reaction and, at equilibrium, the **rate of the forward reaction equals the rate of the reverse reaction**. So if we graph rate versus time for each reaction, the lines will meet and become the same straight line.

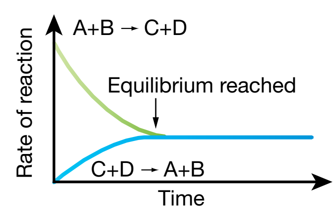


Figure 3.2 Equilibrium and rate of reaction.

3. The **macroscopic properties** (ones you can see) **stay constant** – there is no change in pH, state, colour, temperature or pressure.
4. **Concentrations of all reactants and products stay constant**. Notice that although the concentration at equilibrium stays the same (constant), the concentration of reactants does not have to be equal to the concentration of products.

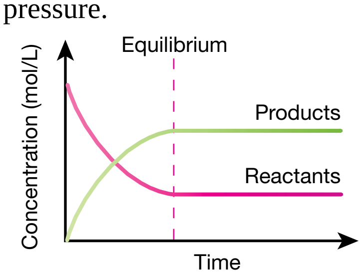
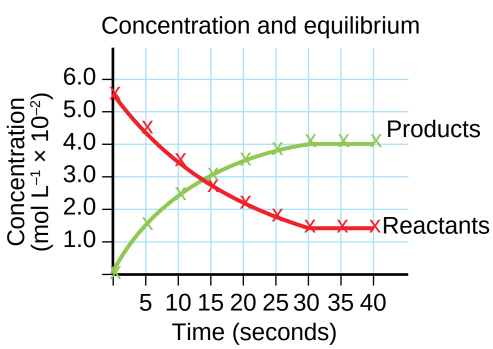


Figure 3.3 Concentration and equilibrium.

5. **Continuous changes** occur at the atomic level, with products forming and breaking up at equal rates (it is a dynamic process) but we cannot see these changes (macroscopic properties are constant).
6. The equilibrium can be approached **from either direction**. For example, in the equilibrium reaction, $A + B \rightleftharpoons C + D$, you can start with the chemicals A and B or with C and D.

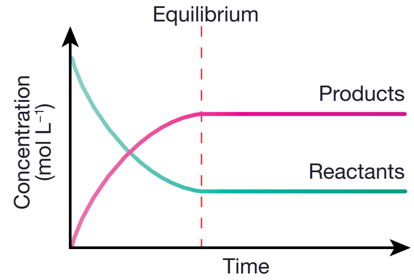
QUESTIONS

1. (a) Outline what is meant by a reversible reaction.
   (b) Identify a macroscopic property you could observe to determine when a reversible reaction reaches equilibrium.
2. Explain why chemical equilibrium systems are described as being dynamic rather than static.
3. Outline the six characteristics of an equilibrium system.
4. Explain how the concentration of reactants can stay constant at equilibrium even though the reaction is still proceeding.
5. (a) Outline trends occurring in the concentration/time graphs in Figure 3.3.
   (b) Explain why both graphs in Figure 3.2 finish as the same straight line.
6. Two colourless solutions contain ferric (Fe^{3+}) ions and thiocyanate (SCN^-) ions. These two ions react when mixed to form a complex ion which is red in colour and has the formula $FeSCN^{2+}$.
   $Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$
   Solutions containing equal concentration of the two ions were mixed and the reaction was allowed to come to equilibrium. The concentration of reactants and products was constantly monitored to produce the graph shown below.

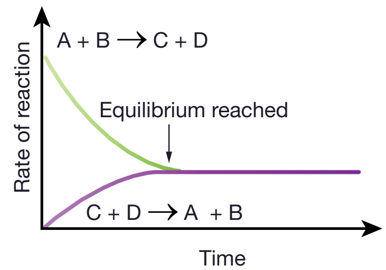


   Identify the time at which equilibrium was reached and justify your answer.
7. Describe an equilibrium system that you observed in the science laboratory.
8. The graphs below illustrate two characteristics of systems at equilibrium.

Graph X

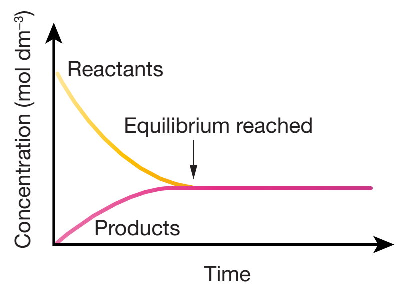


Graph Y

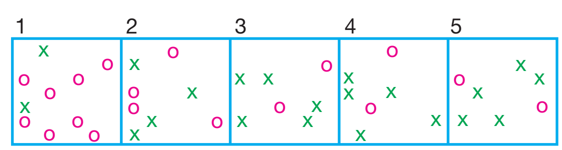


   (a) Outline the characteristic illustrated by each graph.
   (b) Compare graphs X and Y above with graph Z below.

Graph Z



   Does graph Z represent a characteristic of all systems at equilibrium? Justify your answer.
9. In the following series of diagrams, molecules of $NO_2(g)$ (shown as O) are reacting to produce $N_2O_4(g)$ (shown as X). Eventually the two gases reach equilibrium.



   (a) At which stage is equilibrium reached? Justify your answer.
   (b) Is this a physical or a chemical equilibrium? Explain.
10. Check your knowledge with this quick quiz.
   (a) What conditions must be constant for a system to reach a stable equilibrium?
   (b) At equilibrium, what becomes constant?
   (c) At equilibrium, what must be equal?

QCE CHEMISTRY

UNIT 3

TOPIC 7 Chemical Equilibrium Systems

Topic 7 Test

Section A – Multiple Choice (20 marks)

1. A system is in a state of equilibrium when:
   (A) No further reaction occurs.
   (B) The rate of forward reaction equals the rate of reverse reaction.
   (C) The forward reaction has gone to completion and produced a precipitate.
   (D) The reaction has stopped except for insignificant sub-atomic changes in both directions.

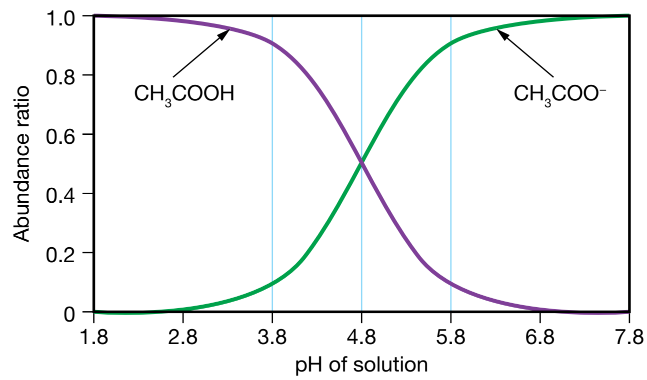
2. An equilibrium position and value will not be shifted by adding:
   (A) More of one reagent.
   (B) More of a product.
   (C) More catalyst.
   (D) Heat.

3. Consider the reaction:
   $2NO(g) + H_2(g) \rightleftharpoons N_2O(g) + H_2O(g) \quad \Delta H = +364$ kJ
   If the temperature is increased and the volume kept constant, the concentration of $N_2O(g)$ will:
   (A) Increase as the forward reaction absorbs energy.
   (B) Decrease as the forward reaction releases energy.
   (C) Remain the same because only the rate of reactions increases.
   (D) Increase as the forward reaction releases energy.

4. In the reactions, $2A + B \rightleftharpoons C + 3D$
   The concentration at equilibrium is:
   A = 0.6 mol L^{-1}, B = 0.5 mol L^{-1}, C = 2 mol L^{-1}, D = 3 mol L^{-1}.
   The equilibrium constant for the reaction is closest to:
   (A) 0.003 (B) 0.18 (C) 54 (D) 300

5. The equilibrium constant for the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is calculated using:
   (A) $\frac{2[NH_3]}{[N_2]\,3[H_2]}$ (B) $\frac{[N_2]\,[H_2]^3}{[NH_3]^2}$
   (C) $\frac{[NH_3]^2}{[N_2]\,[H_2]^3}$ (D) $\frac{[NH_3]}{[N_2]\,[H_2]}$

6. The graph shows how the abundance of ethanoate ions and ethanoic acid molecules changes in an ethanoic acid/ethanoate buffer at different pH values.



   This buffer would be most effective at a pH of 4.8 because, at that pH:
   (A) Most of the ethanoic acid exists as molecules.
   (B) Most of the ethanoic acid exists as ethanoate ions.
   (C) The concentration of acid = the concentration of base.
   (D) The pH > pK_a

7. Consider the following system at equilibrium:
   $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
   Which statement must be correct if the system is at equilibrium?
   (A) The concentrations of H_2, I_2 and HI are equal.
   (B) The concentration of HI is twice that of H_2 and I_2.
   (C) The microscopic properties are constant.
   (D) The rate of the forward reaction equals the rate of the reverse reaction.

Answers

Topic 7 Chemical Equilibrium Systems

1 Open and Closed Systems

1. A physical change – No new substance is produced, particles stay the same and it is usually easy to reverse. Examples include change of state – melting, evaporation, condensation.
   A chemical change – A new substance forms, particles are changed as atoms are rearranged and it is usually difficult to reverse.
2. (a) A system refers to any part of the Universe that is being studied.
   (b) An open system is one that interacts with its environment – energy and matter can move between the system and its environment.
   (c) A closed system does not interact with its environment – energy can move in and out of the system but matter cannot.
   (d) A reversible reaction is one that goes in both directions – forward and in reverse.
3. (a) (i) Iron(III) chloride + water $\rightleftharpoons$ iron(III) hydroxide + hydrochloric acid
   (ii) Copper(II) sulfate (hydrated) $\rightleftharpoons$ anhydrous copper(II) sulfate + water
   (iii) Ammonium chloride $\rightleftharpoons$ ammonia + hydrogen chloride
   (iv) Carbon dioxide + water $\rightleftharpoons$ carbonic acid
   (v) Oxygen $\rightleftharpoons$ ozone
   (b) (i) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
   (ii) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
   (iii) $2NO_2(g) \rightleftharpoons N_2O_4(g)$
   (iv) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$
   (v) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

2 Investigating the Reversibility Of Chemical Reactions

1. (a) Safety precautions for combustion of magnesium and other substances include, wearing goggles to protect the eyes from sparks and burning materials, do not look at the burning material as the light may be very bright and damage the eyes, ensure the burning material is being securely held or supported so it cannot fall and start a fire, ensure that there is no combustible material nearby or underneath the burning substance to avoid starting a fire. A fire blanket and extinguisher should always be on hand in science laboratories in case of accidents.
   (b) Various, e.g. wear goggles to prevent chemicals being splashed into the eyes and always wash hands thoroughly to remove any chemicals as they may be toxic, e.g. ingestion of copper chloride is toxic.
2. (a) To dehydrate pink cobalt chloride crystals the water of crystallisation must be removed. This can occur if the crystals are placed in a desiccator, or if they are heated gently to evaporate the water.

   $CoCl_2.6H_2O(s) \rightleftharpoons CoCl_2(s) + 6H_2O(g)$
   (b) When blue copper sulfate crystals are heated the water of crystallisation present in their crystals evaporates and a white solid (anhydrous copper sulfate) is produced. This reaction can be reversed. If a few drops of water is added to the anhydrous copper sulfate it re-forms the blue crystals of hydrated copper sulfate.
   $CuSO_4.5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(g)$
   (c) Various, for example: Cobalt chloride paper can be used to detect for the presence of water, e.g. to determine whether or not an unknown liquid is water, to detect any water leaking from a pipe or container, in biology experiments to see whether water is released from both top and bottom surfaces of a leaf.
3. A desiccator is a sealable container used to absorb moisture or provide a dry environment. It usually contains a drying agent (desiccant) below the plate with holes. The drying agent is a hygroscopic substance (it absorbs water) such as anhydrous calcium chloride or silica gel. The lid is greased to make an airtight seal. Some desiccators can have a vacuum applied.
4. (a) Repeat the experiment, replacing each ion in turn with a different ion and see if the reaction still happens, e.g. instead of using iron nitrate, you could use calcium nitrate; instead of using potassium thiocyanate, you could use sodium thiocyanate.
   (b) The red colour of blood is due to the haemoglobin molecule which contains iron in a similar structure to the iron complex produced in this reaction. (*Note:* All animals do not have red blood – you might want to find out more about this.)
5. (a) All of the reactants must not have been used up – the reaction had not gone to completion.
   (b) This also shows that the reaction had not gone to completion – adding excess of either reactant can produce more product. Both reactants were still present, along with the product.
   (c) For the colour to become lighter, some of the product must have broken up and re-formed the reactants which were both much lighter in colour. The reaction must have gone in the reverse direction.

3 Characteristics Of Equilibrium

1. (a) A reversible reaction is a reaction which proceeds in both directions at the same time.
   (b) Various. Macroscopic properties include state (solid/liquid/gas), colour, temperature and pressure. For example, you could observe the colour – when the colour stays constant (there are no longer any visible changes taking place) then the system is at equilibrium.
2. Although the macroscopic properties stay constant at equilibrium, there is continual change happening at a submicroscopic level as both the forward and reverse reactions continue to occur. A static equilibrium can be found in physical systems, e.g. riding a bike – where being in equilibrium means that you do not fall off because opposing forces are balanced.
3. The characteristics of a system at equilibrium are:
   - It is a closed system – no matter or energy enters or leaves the system.
   - It involves a reversible reaction and the rate of the forward reaction equals the rate of the reverse reaction.
   - The macroscopic properties (ones you can observe) stay constant – there is no change in state, colour, temperature or pressure.
   - Concentrations of all reactants and all products stay constant.
   - Continuous changes occur at the atomic level, with products forming and breaking up at equal rates, but we cannot see this.
   - The equilibrium can be approached from either direction.
4. As reactants are used up, products are replacing the reactants. Both reactions, the forward and reverse reactions, occur at the same rate, so there is no change in the concentrations of either the reactants or the products.
5. (a) The concentration of reactants decreases, rapidly at first and then less rapidly as reactants are used up. The concentration of products is initially zero, but then product forms increasingly rapidly. Eventually the rates of the two reactions become the same – the rate of the forward reaction equals the rate of the reverse reaction. At this stage – shown by the dotted vertical line on the graph – the system has reached equilibrium. Once a system is at equilibrium, the concentrations of all reactants and products stay constant – they do not change. On the graph, the concentrations while the system is at equilibrium are shown as horizontal lines – indicating that each has a constant value.
   (b) In Figure 3.2, the same straight line at the end of each graph indicates that the rate of the forward reaction is the same as the rate of the reverse reaction, and both are staying constant – they are not changing. This occurs when a system is in equilibrium.