YEAR 12 CHEMISTRY

EQUILIBRIUM AND ACID REACTIONS

LESSON 1: REVERSIBLE REACTIONS

SAMPLE RESOURCES

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MATRIX EDUCATION
THE HSC EXPERTS

1. Equilibrium

Students:
- model static and dynamic equilibrium and analyse the differences between open and closed systems
- investigate the relationship between collision theory and reaction rate in order to analyse chemical equilibrium reactions

❐ Irreversible and Reversible Reactions

- The reactions we have dealt with so far are considered **irreversible**.
  - Changing the **conditions** (e.g. temperature, pressure) won't cause the products to change back to reactants.
  - Irreversible reactions proceed until one of the reactants is **completely used up**.

- For example, the combustion of magnesium in excess oxygen is considered an irreversible reaction.
  - Write a balanced chemical equation for this reaction.

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  - How could you test if the reaction has gone to completion and is irreversible?[1]

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  - Give another example of an irreversible reaction that goes to completion.

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- Although there are many examples of irreversible reactions, some reactions are **reversible** and **do not go to completion**. They are capable of proceeding in both the forward and reverse directions.

- For example, when nitrogen reacts with hydrogen to form ammonia, some ammonia decomposes back into nitrogen and hydrogen.
  - Write a balanced equation for the formation of gaseous ammonia.

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  - Write a balanced equation for the decomposition of gaseous ammonia.

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- For reversible reactions, instead of writing both reactions every time, we use a **reversible arrow** to indicate that both reactions are capable of proceeding:

Reactants ⇌ **Products**

(products of reverse reaction) (reactants of reverse reaction)

  - Write an equation for the reversible reaction producing ammonia.

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- Many chemical changes are reversible. **Physical changes** are generally reversible.
  - For example, the boiling of water is reversible.
  - Write a balanced equation for the boiling of water, using a reversible arrow.

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❒ Static and Dynamic Equilibrium

- Equilibrium refers to the state of a **closed** chemical system in which:
  - the **concentrations** of both reactants and products **do not change** with time
  - the **rate** of the forward reaction is **equal** to the rate of the reverse reaction

- The rate of reaction is how rapidly a reaction proceeds.
  - What does the rate of reaction depend on? (Hint: Think back to collision theory.)

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  - What changes will increase the rate of a reaction?

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  - How can the rate of reaction be measured?[2]

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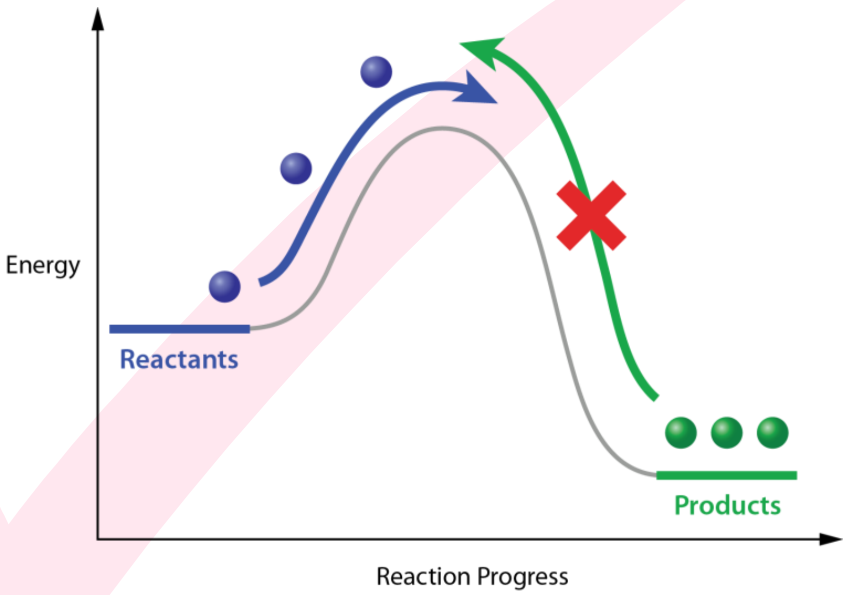
Note to Students

- An open system is a system where matter and energy can enter and leave.
- A closed system is a system where matter cannot enter and leave, but energy exchange can take place with the surroundings (in the form of pressure or heat).

- **Irreversible reactions** (shown with a forward arrow →) that go to completion reach a **static equilibrium**.
  - What are the rates of the forward and reverse reactions when static equilibrium has been achieved?

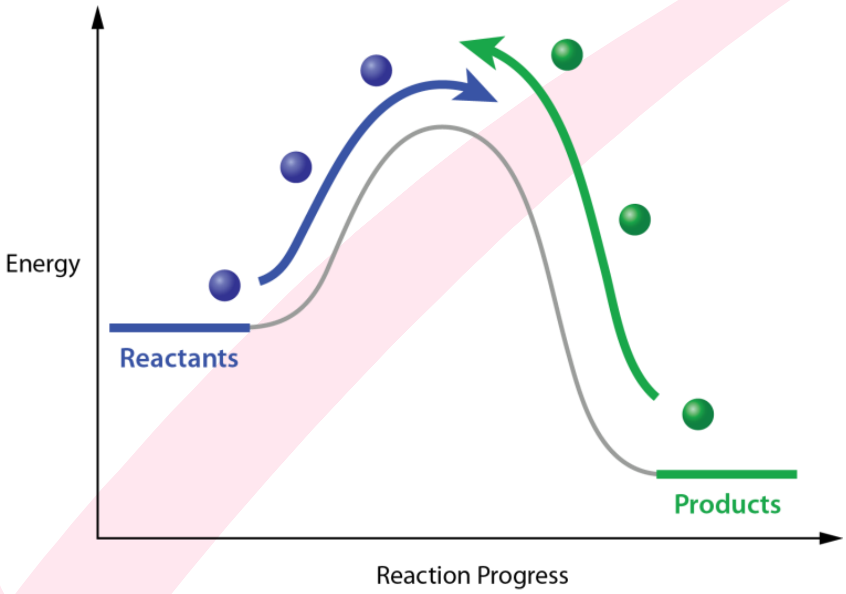
  - Will there be any macroscopic signs of change at equilibrium?

  - The products remain products and do not change back to reactants.



- **Reversible reactions** (shown with a reversible arrow ⇌) do not go to completion.

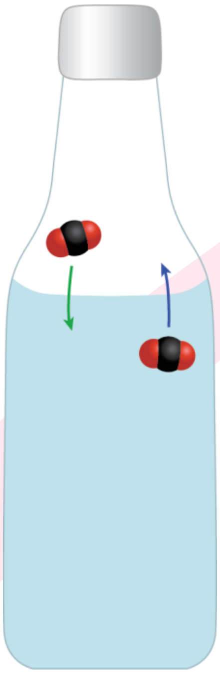
- In a closed system, reversible reactions will instead reach a state known as **dynamic equilibrium**.
  - At equilibrium, the rates of the forward and reverse reactions are the same, but they are non-zero.
  - The equilibrium is dynamic because there are changes occurring at the microscopic level, even though the system undergoes no change at the macroscopic level.



> **Note to Students**
>
> The term "equilibrium" generally refers to dynamic equilibrium.

❒ Analysing Dynamic Equilibrium

- An example of a system that is at dynamic equilibrium is a **sealed bottle of soda water**.
  - There is dissolved carbon dioxide in the liquid, and carbon dioxide in the space above the liquid.
  - If the lid is kept on the bottle, carbon dioxide molecules in the space above the liquid cannot escape from the bottle (**closed system**).



$$CO_{2(g)} \rightleftharpoons CO_{2(aq)}$$

  - If aqueous and gaseous carbon dioxide are in equilibrium with each other, what does this say about the forward and reverse **rates**?

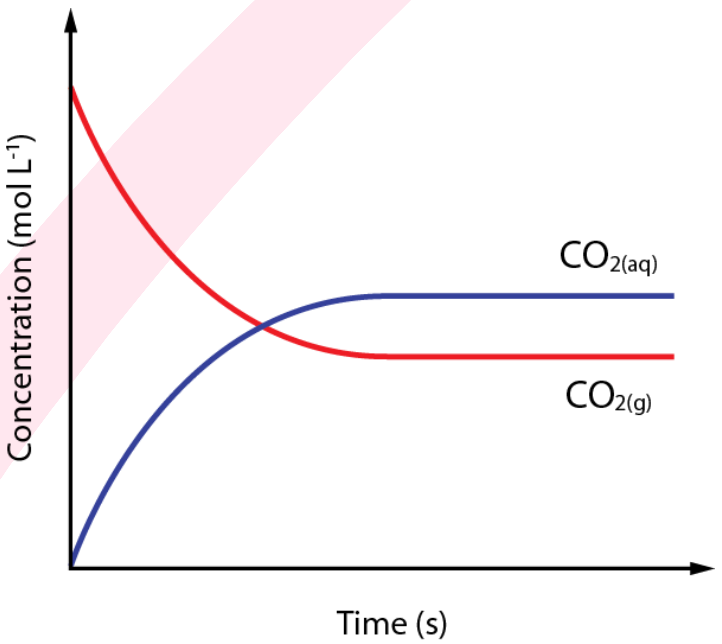
  - Therefore, if a molecule of CO_2 gas in the sealed bottle enters solution and dissolves, what must be occurring simultaneously?

- What effect does this have on the **concentrations** of reactants and products at equilibrium?

- Therefore there are **no observable macroscopic signs of change** in the system at dynamic equilibrium.
  - Microscopically, there are trillions of carbon dioxide molecules "jumping" into and out of solution at any one time.
  - However, the microscopic changes balance out to produce **no overall net change** at the macroscopic level.

- Changes in the concentrations of reactants and products during the course of the reaction can be presented using a **concentration profile diagram**.

$$CO_{2(g)} \rightleftharpoons CO_{2(aq)}$$



  - It shows the **changing concentrations** of reactants and products as the reaction proceeds.
  - **All species are present** in the reaction mixture at equilibrium.

- Equilibrium refers to **equal rates** of the forward and reverse reactions, **not equal concentrations** of reactants and products.

- Refer to the concentration profile diagram to answer the questions below.
  - What is present in the container initially?
  - How do the concentrations of the reaction components change in the first part of the reaction?
  - What is the relationship between concentration and rate?
  - How does the rate of the forward reaction change as the system proceeds towards equilibrium? Why?
  - How does the rate of the reverse reaction change as the system proceeds towards equilibrium? Why?

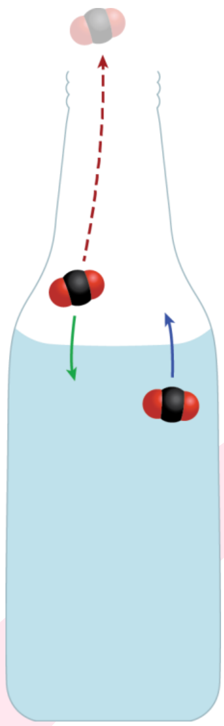
- How do the concentrations of the reaction components change after equilibrium has been reached? Explain.

- Explain why the pressure of the system does not change at equilibrium.

- The forward reaction is exothermic. Explain why the temperature of the system does not change at equilibrium.

- Now consider if the bottle was **left open** to the environment (**open system**), so any CO_2 gas present can escape.

$$CO_{2(g)} \rightleftharpoons CO_{2(aq)}$$



– How would the rate of the forward reaction be affected, compared to in a closed system?

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– How would the rate of the reverse reaction be affected, compared to in a closed system?

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– Would a dynamic equilibrium be achieved? Explain.

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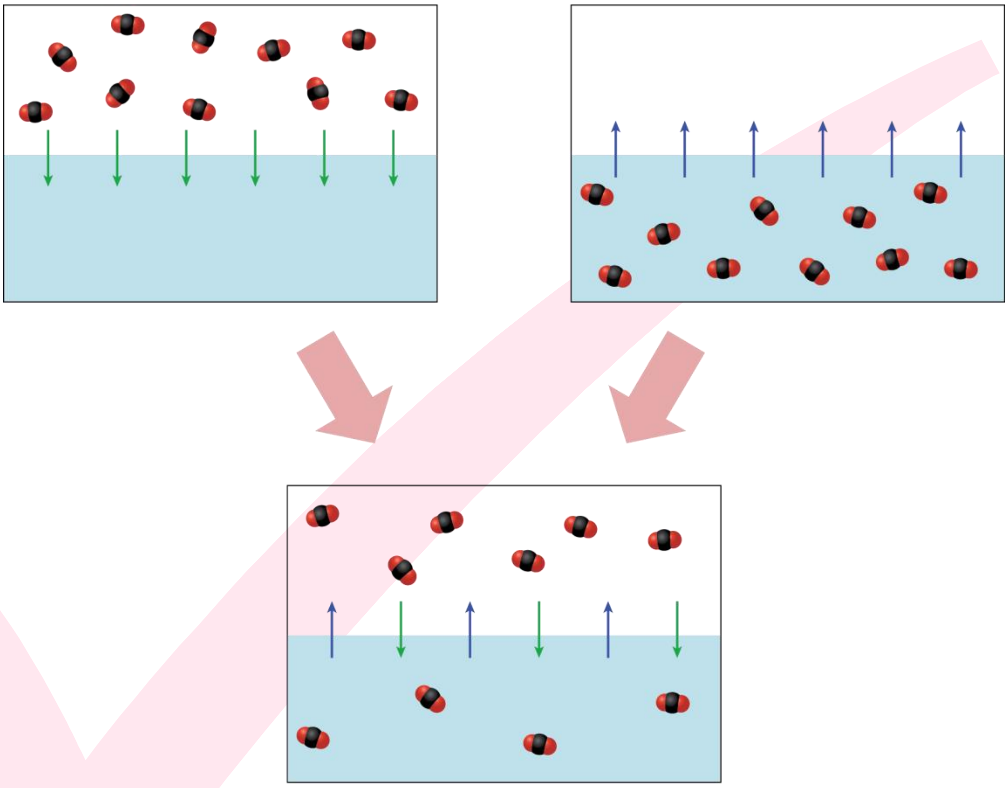
– Hence, equilibrium can only be established in a **closed system**.

- Because the reaction is reversible, the equilibrium **can be approached from either side**.
  - For the CO_2 equilibrium, we can start with all of the carbon dioxide particles as gas, or the same number of carbon dioxide particles as dissolved aqueous molecules.
  - The systems will arrive at the **same equilibrium position** under the same temperature and pressure conditions.


Equilibrium

Concept Check 1.1

A sample of colourless dinitrogen tetroxide decomposes to form brown nitrogen dioxide inside a sealed flask. This process is reversible, and the decomposition reaction is endothermic.

(a) Write a balanced chemical equation for this reaction.[3] **1**

(b) Explain how the rates of the forward and reverse reaction change as the reaction mixture proceeds towards equilibrium.[4] **2**

(c) State three observations that could be made if the system is at equilibrium. Explain how each observation indicates that equilibrium has been established.[5] **3**

Concept Check 1.2

Compare:

(a) open and closed systems[6] **2**

(b) reversible and irreversible reactions[7] **2**

(c) static and dynamic equilibrium[8] **2**

2. Drivers of Reactions

Students:
- analyse examples of non-equilibrium systems in terms of the effect of entropy and enthalpy, for example:
  - combustion reactions
  - photosynthesis

❒ Enthalpy, Entropy and Free Energy

- Recall the two **driving forces** for physical and chemical changes: enthalpy and entropy.

- **Enthalpy (H)** is the heat content of a system.
  - Absolute enthalpy cannot be measured. Instead, the **change in enthalpy (∆H)** is measured.
  - ∆H is a measure of how much **heat energy** is released or absorbed in a reaction or physical process.
  - What units are usually used for ∆H?

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  - What does the sign of ∆H tell you about a reaction?

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  - Sketch an enthalpy diagram for a reaction that absorbs 120 kJ mol^{-1}.

- If the activation energy for the reaction you have drawn above is 400 kJ mol^{-1}, what is the activation energy for the reverse reaction?[17]

- How is ΔH related to bond breaking and bond making?

- How can ΔH be experimentally determined?

- Which law allows you to calculate ΔH indirectly?

- How does enthalpy relate to spontaneity?[18]

- **Entropy (S)** is a measure of the dispersion of available energy. Greater dispersal of energy will increase entropy. It is also sometimes referred to as a measure of disorder.
  - Can absolute entropy be measured?
  - What units are usually used for entropy?
  - What changes will increase entropy?[19]
  - How does entropy relate to spontaneity?[20]

- According to the **second law of thermodynamics**, the entropy of the universe must remain the same or increase for all spontaneous processes.
  - **Gibbs free energy (G)** allows us to compare the relative contributions of enthalpy and entropy to predict overall spontaneity:

$$\Delta G = \Delta H - T\Delta S$$

  - How does Gibbs free energy relate to spontaneity?[21]

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  - Using the equation above, state which of the following will result in a reaction which is:
    - spontaneous at all temperatures
    - non-spontaneous at all temperatures
    - spontaneous at high temperatures
    - spontaneous at low temperatures

		ΔH	
		−	**+**
ΔS	**+**		
	−		

  - Note that a **non-spontaneous reaction can still occur** if it is driven by a spontaneous process, so that the overall entropy of the universe increases. For example, electrolysis will occur if given an input of energy by a battery (galvanic cell).