

Year 12 Chemistry

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Chapter 1

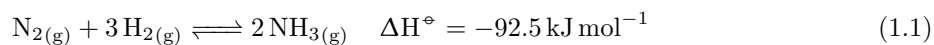
Module 5 Equilibrium and Acid Reactions

1.1 Le Chatelier's Principle

"If a system at equilibrium is subject to a change in conditions, then the system will behave in such a way so as to partially counteract the imposed change"

Haber process:

1.1.1 Effect of Concentration



1.1.2 Effect of Pressure

1.1.3 Effect of Partial Pressure

1.1.4 Effect of Volume

Decreasing the volume will increase the pressure. (Boyle's Law) This increases the collision rate between the reactants and favours the forward reaction.

1.1.5 Effect of Temperature

asd

1.1.6 Summary

To use Le Chatelier's principle to predict the outcome of a change in conditions, you need to consider the following points.

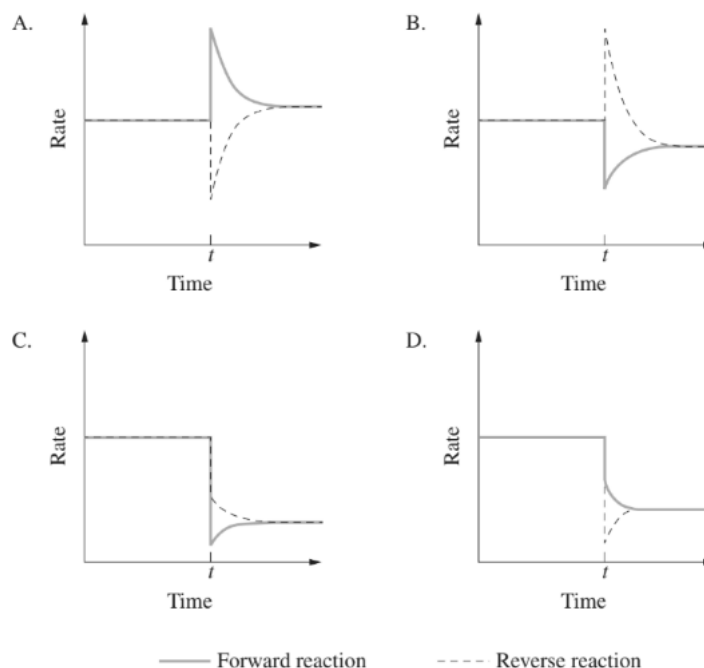
1. What change is imposed?
2. What is the opposite of the change?

3. Which reaction direction is favoured - the forward or reverse?
4. Does equilibrium shift to the left or right?
5. What happens to the concentrations of each aqueous substance or gas?

19 Nitrogen dioxide reacts to form dinitrogen tetroxide in a sealed flask according to the following equation.



Which graph best represents the rates of both the forward and reverse reactions when an equilibrium system containing these gases is cooled at time t ?



D When temperature decreases, the rates of both forward and backward reactions will decrease regardless of which way the endothermic or exothermic reaction goes. (A and B can be eliminated)

This is because all the particles in the system lose kinetic energy, decreasing the rate of collisions hence, decreasing the rate of reaction.

However, since there is a decrease in temperature the exothermic reaction will be favoured in order to counteract the change. In this case, the forward reaction being exothermic is affected less by the drop in temperature as shown in D.

1.2 Practical Investigation 2.3 - Effect of changes to concentration on equilibrium

Aim: To observe the effect of a change in concentration on a system at equilibrium

1.2.1 Materials

- 2 mL of 0.1 mol L^{-1} iron(III) chloride solution
- 2 mL of 0.1 mol L^{-1} ammonium thiocyanate solution
- 1 mL of 0.1 mol L^{-1} calcium fluoride solution
- 20 mL distilled water
- 2x 10 mL measuring cylinders
- 25 mL measuring cylinder
- 4 test tubes
- Test-tube rack
- 4 small labels
- Disposable 1 mL droppers
- Waste bottle
- Digital camera
- Safety glasses

1.2.2 Risk Assessment

Hazard	Precaution
Chemicals may splash onto skin or eyes	Wear safety glasses and wash hands
Chemicals may harm aquatic life	Place in inorganic waste container

1.2.3 Method

1. Pour 1 mL of iron(III) chloride solution into a 10 mL measuring cylinder.
2. Pour 1 mL of ammonium thiocyanate into another 10 mL measuring cylinder.
3. Pour both solutions into the 25 mL measuring cylinder.
4. Add 18 mL of distilled water to the 25 mL measuring cylinder so that the total volume is 20 mL.
5. Label four test tubes A, B, C and D.
6. Pour equal volumes of the solution in the 25 mL measuring cylinder into each of the test tubes.
7. Retain test tube A as the reference solution.
8. Add 1 mL of iron(III) chloride to test tube B.
9. Take a photo to record observations for test tube B relative to test tube A.
10. Add 1 mL of ammonium thiocyanate to test tube C.

11. Take a photo to record observations for test tube C relative to test tube A.
12. Add 1 mL of calcium fluoride to test tube D. (Note: This reacts with the iron(III) ion so there is less iron(III) available to react with the thiocyanate ion.)
13. Take a photo to record observations for test tube D relative to test tube A

1.2.4 Results

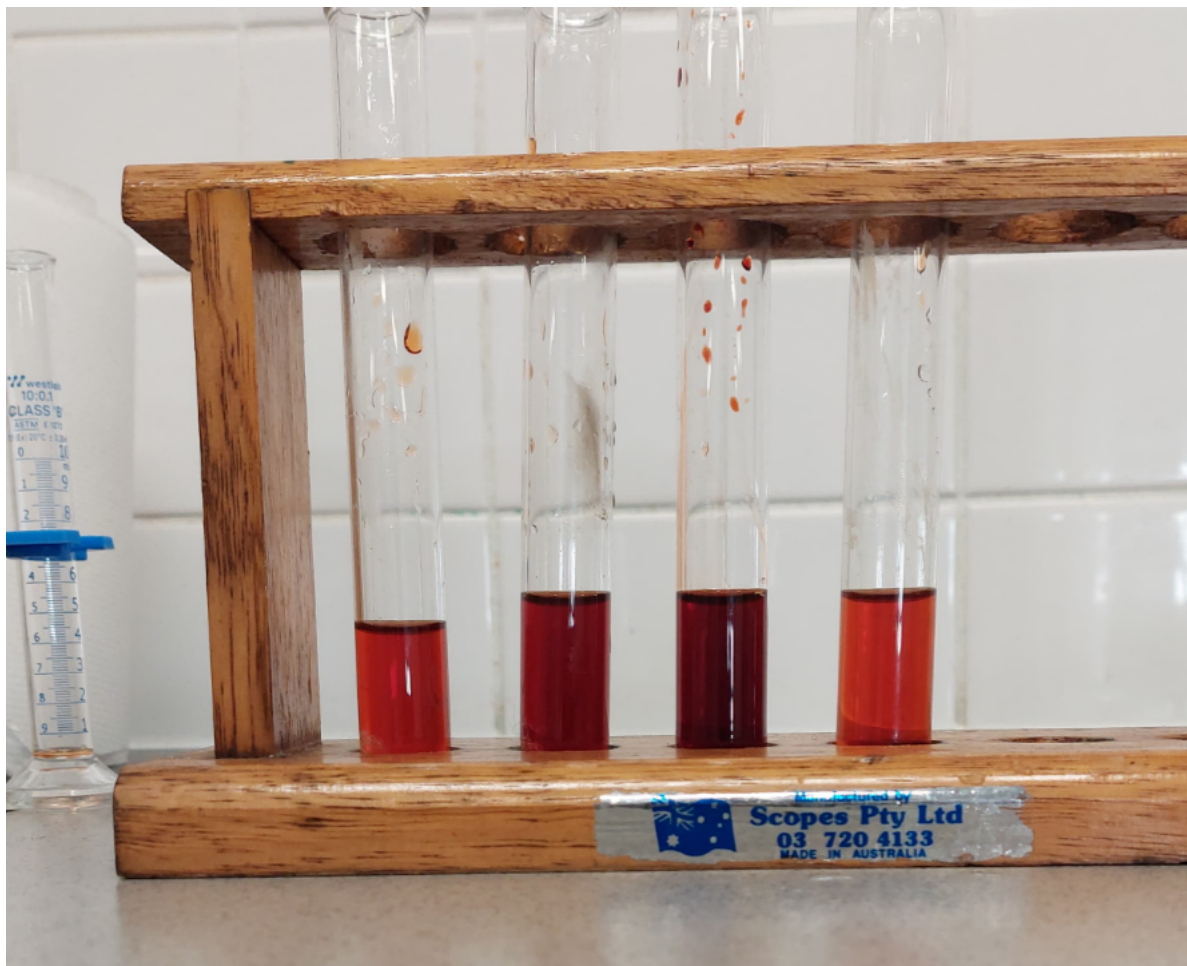


Figure 1.1: Test tubes A, B, C, D

1.2.5 Discussion

Explain each colour change in terms of collision theory.

The test tube B was darker in colour in comparison to test tube A. The increase in moles of reactants allows more successful collisions to occur, increasing the amount of product. The same principle applies to test tube C.

Test tube D was lighter in colour compared to A, due to the calcium fluoride reacting with the iron (III) chloride

1.2.6 Conclusion

Use Le Chatelier's principle to explain what happened in test tubes B, C and D.

Test tube B was darker due to the increase in concentration of the reactant iron (III) chloride causes a shift of the equilibrium towards the products due Le Chatelier's principle

Test tube C was darker due to the increase in concentration of the reactant ammonium thiocyanate causes a shift of the equilibrium towards the products due Le Chatelier's principle

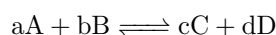
Test tube D was lighter because the calcium fluoride reacted with the iron (III) chloride, lowering the overall concentration of iron (III) chloride. This reduced the amount of reactants available, making the reverse reaction more favourable by Le Chatelier's principle.

1.3 Calculating the Equilibrium Constant

The equilibrium constant can be used to predict the direction of chemical reactions

$$K_{eq} = \frac{[products]}{[reactants]}$$

For reaction:



the equilibrium expression is:

$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The concentration of each chemical species is raised to the power of the number of moles of that species indicated in the chemical equation. (Eg. there are d moles of species D , hence in the equilibrium expression the concentration of species D is raised to the power of d , written as $[D]^d$)

- The value for the equilibrium constant only takes into account the concentration of substances where the concentration can vary
- Solutions and gases can vary in concentration or partial pressure hence are included in K_{eq}
- Solids and pure liquids are NOT included eg. H_2O isn't required when calculating

1.3.1 Reaction Quotient (Q)

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Has the same formula as K_{eq} , however applies to any stage of a reaction

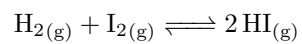
1.3.2 Equilibrium Constant

- Comparing Q to K_{eq} predicts which way the equilibrium will shift
- K_{eq} is where the equilibrium lies
- A large K_{eq} means that there are more products than reactants; ie. equilibrium lies towards completion
- If K_{eq} is close to one, both reactants and products are plentiful at equilibrium

Example Let $Q = 2.1$, $K_{eq} = 0.315$, \therefore products > reactants.

1.3.3 Calculating the equilibrium expression

Consider the reaction between hydrogen and iodine producing hydrogen iodide:



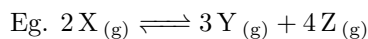
1.4 ICE Tables

	[A]	[B]	[C]
Initial concentration			
Change in concentration			
Equilibrium concentration			

$$\Delta c = c_{eq} - c_u$$

Eg. $\text{A} + \text{B} \rightleftharpoons \text{C} + \text{D}$

	[A]	[B]	[C]	[D]
Initial concentration	0.6	0.6	0	0
Change in concentration	-0.5	-0.5	+0.5	+0.5
Equilibrium concentration	0.1	0.1	0.5	0.5



A sample consisting of 0.500 mol of X is placed into a system with a volume of 0.750 liters. At equilibrium, the amount of sample X is known to be 0.350 mol.

	X	Y	Z
I	0.5	0	0
C			
E	0.35		

	X	Y	Z
I	0.5	0	0
C	-0.15	-0.225	+0.3
E	0.35	0.225	0.3

$$[X] = \frac{0.35}{0.75} = 0.467$$

$$[Y] = \frac{0.225}{0.75} = 0.3$$

$$[Z] = \frac{0.3}{0.75} = 0.4$$

1.5 Effect of Temperature on K_{eq}

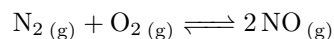
Although other factors may affect equilibrium, K_{eq} is only affected by temperature. Changing concentration, pressure, or volume will change the concentrations and therefore adjust the reaction point, however the reaction will still equalise to achieve the same K_{eq}

- For a particular reaction, K_{eq} is constant at a given temperature
- Temperature changes the ratio of products and reactants, hence changing K_{eq}
- For $\text{N}_2\text{O}_2_{(g)} \rightleftharpoons 2\text{NO}_{2(g)}$, temperature increases the K_{eq} value and favours the formation of products. The forward reaction is endothermic

1.5.1 Example Question

Nitric oxide gas (NO) can be produced from the direct combination of nitrogen gas and oxygen gas in a reversible reaction.

1. Write a balanced chemical equation for this reaction (1 mark)



2. Explain, using collision theory, how an increase in temperature would affect the value for K_{eq} for this system. Refer to the diagram in your answer.

An increase in temperature would favour the forward reaction, hence K_{eq} will increase. More energy allows more collisions to occur

$$K_{eq} = \frac{[p_1][p_2]}{[r_1][r_2]}$$

$$1 = \frac{[1][1]}{[1][1]}$$

If p_2 decreases to $[0.5]$, favouring the forward reaction

$$\begin{aligned} &= \frac{[1.15][0.65]}{[0.85][0.85]} \\ &= K_{eq} \end{aligned}$$