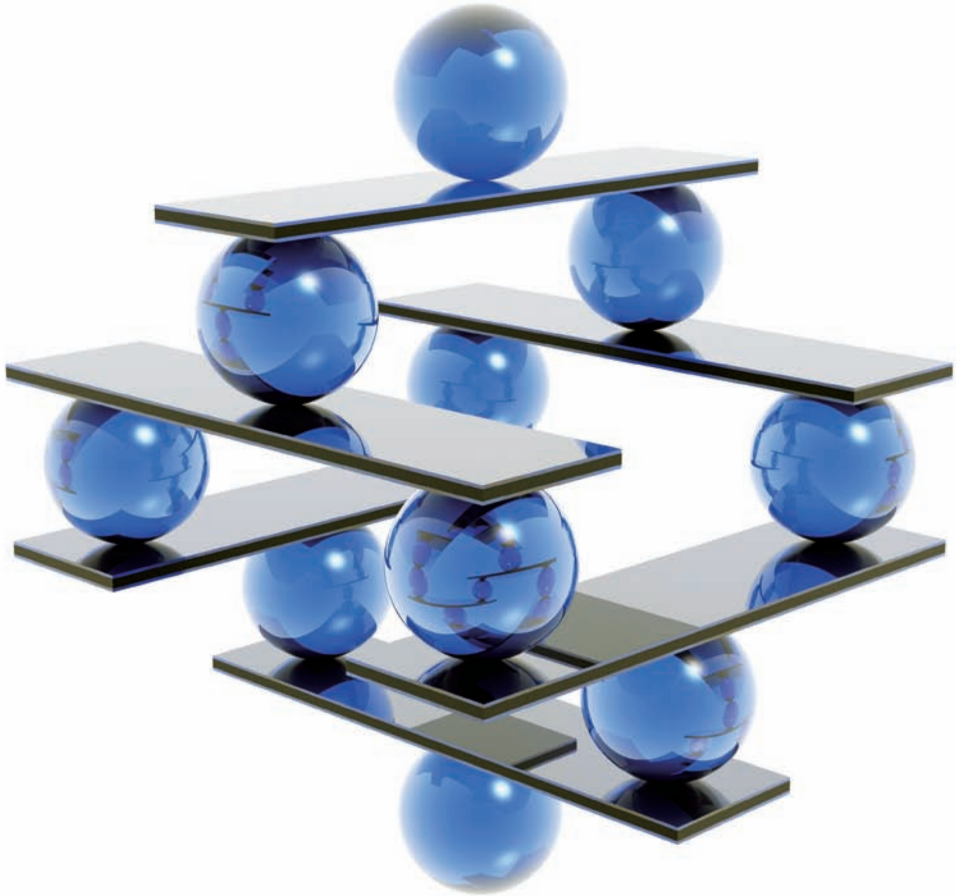
Equilibria

Student Notes



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Activity 1: Heating ammonium chloride

Introduction

In some chemical reactions, the products can react to form the original reactants. These are called reversible reactions.

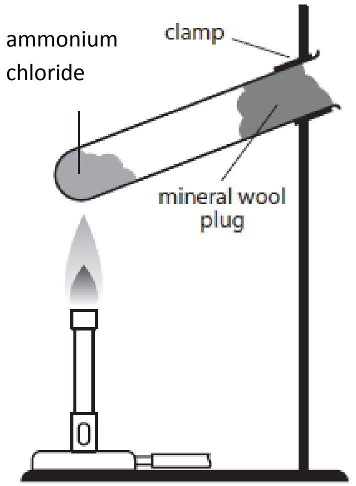
For example, ammonium chloride is a white solid that breaks down into ammonia and hydrogen chloride gases when heated.

These gases react to form ammonium chloride

What to record

Record any observations.

What to do



1. Set up the apparatus as shown.
2. Heat the ammonium chloride.
3. Allow to cool.

Safety

Wear eye protection.

Questions

1. What is the white solid that is formed on cooling?
2. What do the following words mean (*a*) decomposition, (*b*) reversible, (*c*) product, (*d*) reaction?
3. Write a word equation for the reaction you have observed.

Activity 5: The equilibrium of the cobalt chloride–water system

In this experiment you will be looking at the reversible reaction shown in the equation below. The hexaquacobalt(II) complex is pink while the tetrachlorocobalt(II) complex is blue. The different colours of the reactants and products allow you to observe the effect on the position of equilibrium of the addition of water and chloride ions and the removal of water.

$$\underset{\text{Pink}}{Co(H_2O)_6^{2+}} + 4Cl^- \rightleftharpoons \underset{\text{Blue}}{CoCl_4^{2-}} + 6H_2O$$

Apparatus (per group)

Safety goggles must be worn

- One plastic well-plate (24 wells) – eg Sigma ref: M9655.

Chemicals (per group)

- Solutions contained in plastic pipettes, see overleaf
- Cobalt chloride (aqueous solution) 0.1 mol dm^{-3} (**Toxic**) See CLEAPSS Hazcard 25
- Cobalt chloride (ethanol solution) 0.1 mol dm^{-3} (Highly Flammable) See CLEAPSS Hazcard 40A
- Concentrated hydrochloric acid (**Corrosive**) See CLEAPSS Hazcard 47A
- Concentrated sulphuric acid (**Corrosive**) See CLEAPSS Hazcard 98A
- Potassium chloride 0.1 mol dm^{-3}
- Deionised water
- Potassium chloride powder. See CLEAPSS Recipe book sheet 34

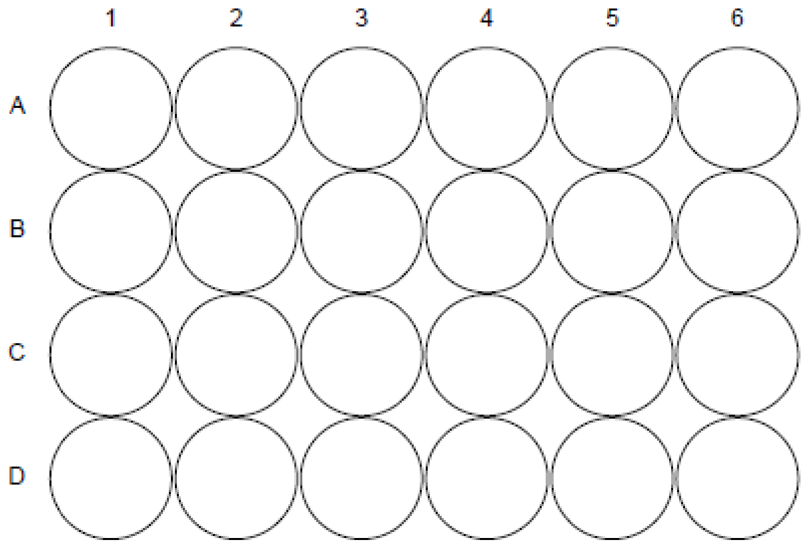
Instructions

Follow the instructions carefully recording the colours of the solutions at each stage. Appropriate care should be taken when using concentrated acids.

1. Put 50 drops of ethanolic cobalt chloride solution into each of the wells A1–A6 and B1 (see diagram below). Place 20 drops of aqueous cobalt chloride into wells B2–B4.
2. Add eight drops of water to wells A2–A5 and carefully swirl the well-plate for 1 min.
3. Add eight drops of potassium chloride solution to wells A6 and B1. Swirl the well-plate gently for 1 min.
4. Add 10 drops of concentrated hydrochloric acid to well A3 and enough solid potassium chloride to wells A4 and A5 to cover the bottoms of the wells. Swirl carefully for 1 min.
5. Add 20 drops, five drops at a time with gentle swirling, of concentrated sulphuric acid to wells A5–B1.
6. Add 20 drops of concentrated hydrochloric acid to wells B3 and B4, swirling gently.
7. Finally, add 30 drops of water to well B4.

Questions

1. Can you give explanations for all your observations – writing equations where appropriate?



Activity 6: The equilibrium of the cobalt chloride–water system; temperature

Introduction

In *SSERC Bulletin 219* [1] we gave details of a simple demonstration showing the effect of temperature change on the position of an equilibrium. Using the solution of cobalt chloride and additional chloride ions the colour change between blue and pink takes place over a particular, smallish temperature range with an intermediate colour of mauve.

When the solution is blue the predominant species is the tetrachlorocobaltate anion with very little hexa-aquocation present. When it is pink the latter cation is the most populous species. The mauve solution contains both of the coloured ions in approximately equimolar proportions. An easy way of convincing pupils of this is for them to hold a *blue* test-tube across a *pink* and see the mauve colour in the area of overlap (Figure 4).

By altering the proportions of cobalt salt and extra chloride added, the solution can be tuned to change colour at different temperatures. A set of tubes so tuned can function as a crude thermometer.

What you will need

Chemicals

cobalt(II) chloride
industrial methylated spirits (IMS, clear)
distilled or de-ionised water
hydrochloric acid (concentrated)
sodium chloride

Equipment

balance
spatula
weighing boat
test tubes, 5 off
pipette, 5 cm^3 or small measuring cylinder
measuring cylinder, 100 cm^3
burette + stand
beakers, 250 cm^3, 6 off
supply of hot water (also iced water optional)
thermometer, 0-100°C

Preparation of solutions

Dissolve 2 g of cobalt chloride in 100 cm^3 of IMS and 16 cm^3 of distilled water.

Tuning each tubeful - Dispense three aliquots of 5 cm^3 of the cobalt solution into 3 test tubes labelled *A*, *B* and *C*. Place one tube in a water bath (250 cm^3 beaker three quarters filled with water at 20°C). Run in concentrated hydrochloric acid from the burette, dropwise with shaking or stirring and allowing time for temperature equalisation, until the colour **just** turns blue. That tubeful is now *tuned* to change colour slightly below 20°C. Repeat with tubes *B* and *C* in the beaker waterbaths at other temperatures, say 30°C and 40°C.

The equilibrium equation can be expressed as follows:

$$[Co(H_2O)_6]^{2+}(aq) + 4Cl^-(aq) \rightleftharpoons CoCl_4^{2-}(aq) + 6H_2O(l) \quad \Delta H \text{ +ve}$$

Pink — Blue (see colours below)

Now place the three tuned tubes in turn in 5 beaker waterbaths and see if the colour change is remembered. Alternatively place them together in the same waterbath, raise the temperature slowly and observe the colour in the tubes turn blue in succession at their tuned temperature. Could this be the next executive desk toy?

Curricular references

Higher Chemistry, Unit 3, Chemical Reactions, (c) - the concept of dynamic equilibrium and shifting the equilibrium position.

Advanced Higher Chemistry, Unit 2: Principles of Chemical Reactions, (b) Chemical equilibrium.

Figure 1 - *A* is tuned to change colour at slightly below 20°C

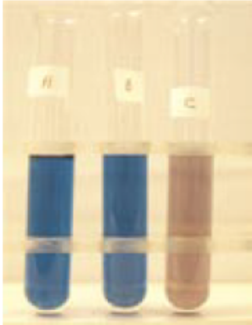


Figure 2 - *B* is tuned to change colour at 30°C

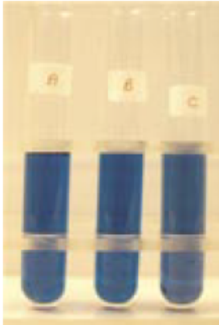


Figure 3 - *C* is tuned to change colour at 40°C

Figure 4 - mauve colour

Reference

1. *SSERC Bulletin No. 219*, Autumn 2006, p6

Chemical	Main Hazard	Control Measures
Cobalt(II) chloride	**Category 2 carcinogen** by inhalation. Sensitiser by skin contact.	Avoid raising dust. Wear eye protection and gloves to prepare solution from the powder. The solution poses negligible risk.
Hydrochloric acid (concentrated)	Extremely irritant and corrosive vapour. Liquid and vapour causes severe burns to eyes, lungs and skin.	Wear nitrile gloves/gauntlets and eye protection. Fuming hydrochloric acid should only be handled in a fume cupboard.

Variations

1. The chloride could be supplied by using saturated sodium chloride solution instead of the acid. This avoids the corrosiveness of the acid. However, owing to the limited solubility of the salt, a larger volume of up to 10 cm^3 is needed to supply a sufficiently high concentration of chloride ions.

2. Using a small scale as described above gives the advantage of a more rapid temperature equilibration. A teacher demonstration might need a larger scale.

3. Ideas for a further extension would be to use it as the basis of an investigation on a more quantitative basis. Using a colorimeter, the concentration of each species could be measured and thus the constancy of the equilibrium constant at a given temperature. If the equilibrium constant were measured at a few temperatures the enthalpy of the reaction could be calculated.

Activity 7: Consolidation

Worksheet

Read each of the statements below.

Tick one of the boxes to show that you think the statement is true, false or that you are unsure whether it is true or not.

When a reactant is added to a system in equilibrium, the forward reaction will occur to use up all the added material and so restore the equilibrium.

When a reactant is added to a system in equilibrium, more product is produced but the value of the equilibrium constant, *K*, remains unaltered.

A system reaches equilibrium when the concentration of the reactants is equal to the concentration of the products.

A high value of the equilibrium constant, *K*, means that the forward reaction is very fast.

When the reactants of an equilibrium system are mixed together the rate of the forward reaction increases until equilibrium is established.

A catalyst increases the rate of reaction of both the forward and reverse reactions.

In an equilibrium system in which the forward reaction is endothermic, the reverse reaction is exothermic.

If a system in equilibrium where the forward reaction is endothermic is heated, the rate of the forward reaction increases but the rate of the reverse reaction decreases.

Statement	The statement is true	The statement is false	I am unsure
1			
2			
3			
4			
5			
6			
7			
8			

For each statement that you think is false, explain why you think it is wrong.

Write down any changes in your thinking as a result your discussions.

Activity 8: Diagnostic test

1 What is Le Chatelier's principle? **(1)**

2 When an equal quantity of iron(II) sulfate solution is added to silver nitrate solution of the same concentration, an equilibrium mixture is formed and a precipitate of silver is seen. The ionic equation is:

$Fe^{2+}(aq) + Ag^{+}(aq) \rightleftharpoons Ag(s) + Fe^{3+}(aq)$

a) What would happen if the concentration of the iron(II) sulfate solution was increased? **(2)**

b) What would you expect to see if iron(III) sulfate solution was added to the equilibrium mixture? **(2)**

3 Using an example for each, explain the difference between reactions involving:

a) thermal dissociation; **(2)**

b) thermal decomposition. **(2)**

4 In the Haber process, ammonia is made in an exothermic reaction between hydrogen and nitrogen gases:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ mol}^{-1}$

Explain, with reasons, what the effect would be on the proportion of ammonia in the equilibrium mixture if:

a) the temperature was increased; **(3)**

b) the pressure on the system was increased. **(3)**

5 a) Write an equation for the reaction between ethanol and ethanoic acid. **(2)**

b) Write an expression for the equilibrium constant for this reaction. **(3)**

Total Mark = 20

Glossary

Catalyst	A substance that increases the rate of a reaction.
Dynamic equilibrium	An equilibrium where the rate of the forward reaction equals the rate of the reverse reaction.
Endothermic	Endothermic reactions absorb heat from the environment.
Equilibrium	At equilibrium the concentration of reactants and products is constant.
Exothermic	Exothermic reactions release energy to the environment.
Irreversible reaction (non-reversible)	Reactions that cannot be reversed by simple means.
Le Chatelier's principle	The position of equilibrium shifts to oppose any change imposed upon it.
Reversible reaction	Changes that can go forwards or backwards.
Thermal decomposition	A substance that is split by heat that is irreversible.
Thermal dissociation	A reaction that can be reversed by altering the temperature.

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