

Experiment 3: LCP and Gases: Injecting gases into and changing temperature of a gas system.

Experiment 4: LCP and Aqueous Solutions:

Experiment 3: LCP and Gases

Random error:

- Inability to adequately identify colour changes.
- Irregular changes in temperature, pressure and humidity.

Systematic error:

- Inherent error in the thermometer.
- Inherent error in the syringe.
- Inherent error in the test tubes.

Experiment 4: LCP and Aqueous Solutions

Random error:

- Inability to adequately distinguish colour changes.
- Parallax error.
- Inability to adequately identify colour changes.

Systematic error:

- Inherent error in measuring cylinders.
- Inherent error in the test tubes.
- Inherent error in the CoCl_2 paper.

Safety note:

- Cobalt chloride is toxic and a possible carcinogen. The indicator papers should be handled with forceps or gloves and hands washed after use.

Experiment 12: Preparation of Standard Sodium Carbonate

Random error:

- Irregular changes in temperature, pressure and humidity.
- Parallax error.

Systematic error:

- Inherent error in the scale.
- Inherent error in the volumetric flask.
- Inherent error in the transfer pipette.

Experiment 13: Preparation and Standardisation of HCl

Random error:

- Parallax error.
- Irregular changes in temperature, pressure and humidity.
- Inability to adequately identify the endpoint colour change.

Systematic error:

- Inherent error in the burette.
- Inherent error in the pipette.
- Inherent error in the measuring cylinders.

Safety note:

- Concentrated HCl is very corrosive and must be handled with extreme care.
You must wear eye protection.
- If any concentrated HCl gets in contact with your skin, immediately wash it off with copious quantities of water.

Experiment 14: Preparation and Standardisation of NaOH

Random error:

- Parallax error.
- Irregular changes in temperature, pressure and humidity.

Systematic error:

- Inherent error in the volumetric flask.
- Inherent error in the scale.
- Inherent error in the pipette.

Safety note:

- NaOH pellets are very corrosive and must not be allowed to come in contact with your skin. If NaOH comes in contact with your skin, wash with copious amounts of water.
- Use a spatula or plastic spoon to handle NaOH pellets.

Experiment 15: Acetic Acid Content in Vinegar

Random error:

- Parallax error.
- Estimating values between graduations.
- Inability to adequately identify the endpoint.

Systematic error:

- Inherent error in the volumetric flask.
- Inherent error in the burette.
- Inherent error in the pipette.

Experiment 31: Reactivity of Alcohols

Random error:

- Irregular changes in temperature, pressure and humidity.
- Parallax error.

Systematic error:

- Inherent error in the test tubes.
- Inherent error in the measuring cylinders.

Safety note:

- Alcohols are flammable liquids and must be kept clear of naked flames. Make sure your Bunsen burner and all nearby Bunsen burners are turned off before you start using the alcohols.
- 6mol/L sulfuric acid is corrosive and must be handled with care. If any H_2SO_4 comes into contact with your skin, immediately wash the affected area with plenty of water.
- Sodium is a very reactive metal, particularly when it comes in contact with water. It must be handled with great care. Wear safety glasses. Use a spatula, plastic spoon or tweezers to handle the sodium. At the end of the experiment, the residual sodium must be disposed of carefully by pouring the alcohol-sodium mixture into a beaker.

Experiment 32: Esters

Random error:

- Irregular changes in temperature, pressure and humidity.
- Parallax error.

Systematic error:

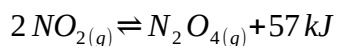
- Inherent error in the test tubes.
- Inherent error in the measuring cylinders.
- Inherent error in the beaker.

Safety note:

- Concentrated ethanoic acid and sulfuric acid are very corrosive and must be handled with extreme care.
- If any of the concentrated ethanoic acid or sulfuric acids comes in contact with your skin, immediately wash the affected area with large quantities of water.
- Most organic substances are flammable and should be kept clear of naked flames.
- Turn off the Bunsen burner before heating the mixtures in a hot water bath to make esters.

Random error	Systematic error:
Inability to adequately distinguish colours.	Inherent error in the thermometer.
Irregular changes in temperature and pressure.	Inherent error in the syringe.
	Inherent error in the test tube

Experiment 3: Le Châtelier's Principle and Gases



Colourless

Equipment:

- Large, transparent glass syringe with sealed end, containing NO₂/N₂O₄.
- 2 sealed tubes or flasks containing NO₂/N₂O₄ at the same pressure.
- Ice bath at around 0°C.
- Hot water bath at around 90°C.
- Thermometer.

- Test tube rack.
- Test tube holder/tongs.
- Bunsen burner.
- Bench protector.
- Matches.

Part A: Effect of pressure or volume on equilibrium.

1. Carefully observe the colour intensity of the gas mixture in the syringe.
2. Push in the gas syringe quickly so that the volume occupied by the gas mixture is instantaneously reduced. Note the immediate change in colour of the gas mixture.
3. Carefully observe the change in the colour intensity which occurs over several seconds after the immediate colour change.

Part B: Effect of temperature on equilibrium.

1. Compare the colours of the 2 flasks containing the $\text{NO}_2/\text{N}_2\text{O}_4$ gas mixtures.
2. Place one of the flasks in an ice bath for 5 minutes and compare the colour in this flask with that in the flask at room temperature.
3. Remove the flask from the ice bath and place it in a hot water bath.
Compare the colour in this flask with that in the flask at room temperature.

Q: In Part A, what happened to the concentration of the 2 gases immediately after the volume was reduced?

It increased.

Q: Explain the immediate change of colour that occurred when the volume of the system was reduced.

Decreasing volume of the system decreases the distance between all particles, increasing the frequency of collisions, increasing the rate of both the forward and reverse reaction. As the molar coefficient of NO_2 is greater than that of N_2O_4 , the pressure of NO_2 will increase more and so the system will appear more brown.

Q: From the colour changed that occurred after the volume was reduced, identify the direction in which the net reaction took place as the system re-established equilibrium.

The reaction turned from brown to a lighter brown. This means that the forward reaction was favoured and so the net reaction is the forward reaction.

Q: Describe your observations using Le Châtelier's Principle.

According to LCP, the system will act to partially counteract the imposed change. As such, increasing pressure by decreasing volume favours the forward reaction which uses up the most particles and hence reduces pressure, and the equilibrium shifts right until a new equilibrium is established.

Q: If the volume of the $\text{NO}_2/\text{N}_2\text{O}_4$ system was increased, predict the effect this would have on the equilibrium and hence the colour intensity in the syringe.

According to LCP, the system will act to partially counteract the imposed change. As such, decreasing pressure by increasing volume favours the reverse reaction which produces the most particles and hence increases pressure, and the equilibrium shifts left until a new equilibrium is established. This means the system will become a lighter brown immediately after the imposed change, and then become a darker brown (not as dark as initially) afterwards.

Q: Given the reaction is exothermic, explain your observations made in Part B.

Increasing temperature increases the average kinetic energy of all particles, increasing the rate of both the forward and reverse reaction. The rate of the reverse, endothermic reaction increases more. This increases the rate of the reverse reaction relative to the forward reaction, and the equilibrium shifts left until a new equilibrium is established. This increases $[\text{NO}_2]$, and so the system becomes a darker brown.

Decreasing

Q: Explain, with reference to collision theory, the effect of changes in concentration or partial pressure on the system.

Increasing pressure decreases the distance between all particles, increasing frequency of successful collisions, increasing the rate of both the forward and reverse reaction. The rate of the forward reaction, which uses up the most particles, increases more. This increases the rate of the forward reaction relative to the reverse reaction, and the equilibrium shifts right until a new equilibrium is established.

Decreasing pressure increases the distance between all particles, decreasing frequency of successful collisions, decreasing the rate of both the forward and reverse reaction. The rate of the forward reaction, which uses up the most particles and hence decreases pressure, decreases more. This decreases the rate of forward reaction relative to the reverse reaction, and the equilibrium shifts left until a new equilibrium is established.

Q: Predict the effect of injecting argon into the $\text{NO}_2/\text{N}_2\text{O}_4$ gas mixture at fixed volume. Justify your prediction with reference to LCP, and explain your prediction with reference to collision theory.

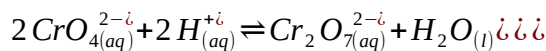
According to LCP, the system will act to partially counteract the imposed change. As such, increasing pressure will favour the rate of the forward reaction which uses up more particles, and the equilibrium shifts right until a new equilibrium is established.

Increasing pressure decreases the distance between all particles, increasing frequency of collisions, increasing the rate of both the forward and reverse reaction. The rate of the forward reaction, which uses up the most particles, increases more. This increases the rate of the forward reaction relative to the reverse reaction, and the equilibrium shifts right until a new equilibrium is established.

Experiment 4: Le Châtelier's Principle and Aqueous Solutions

The 2 reactions being investigated:

1. Chromate-dichromate equilibrium.



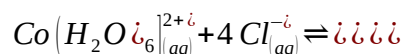
Yellow

Orange

In the KCrO_4 solution $[\text{CrO}_4^{2-}]$ is relatively large and $[\text{Cr}_2\text{O}_7^{2-}]$ is relatively small. In the $\text{K}_2\text{Cr}_2\text{O}_7$ solution $[\text{CrO}_4^{2-}]$ is relatively small and $[\text{Cr}_2\text{O}_7^{2-}]$ is relatively large.

2. The equilibrium between $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{CoCl}_4]^{2-}$.

When $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ is dissolved in water, it forms a pink solution because of the hydrated cobalt ion ($[\text{Co}(\text{H}_2\text{O})_6]^{2+}$). If the $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ is dissolved in concentrated HCl , it forms a deep blue solution because of the formation of the $[\text{CoCl}_4]^{2-}$ ions.



Pink

Blue

Equipment:

- 4 strips of CoCl_2 paper.
- 0.1mol/L $\text{K}_2\text{Cr}_2\text{O}_7$ solution (5mL).
- 0.1mol/L KCrO_4 solution (5mL).
- 0.2mol/L NaOH solution (5mL).
- 0.2mol/L HCl solution (5mL).
- 1mol/L HCl solution (5mL).
- Distilled water.
- 4 large test tubes.
- Graduated cylinder.
- Beaker.
- Dropper.
- Gloves.

Preparation of CoCl_2 paper:

1. Dissolve 5g of hydrated CoCl_2 in 100mL of water (the solution is toxic).
2. Soak the filter paper in this solution and drain and dry in an oven so that it's a definite blue colour.
3. The filter paper can be cut into small strips and stored in a desiccator with dry silica gel.

Part A: The chromate-dichromate equilibrium.

1. Place around 1mL of 0.1mol/L K_2CrO_4 solution into each of 2 test tubes and note the colour of the solution. Use one for comparison of colours.
2. Place around 1mL of 0.1mol/L $\text{K}_2\text{Cr}_2\text{O}_7$ solution into each of 2 test tubes and note the colour of the solution. Use one for comparison of colours.
3. To one of the test tubes of K_2CrO_4 solution, add 0.2mol/L HCl solution dropwise until a colour change is noted.
4. To the same solution, now add 0.2mol/L NaOH solution dropwise until another colour change is observed. Record your observations.
5. To one of the test tubes of $\text{K}_2\text{Cr}_2\text{O}_7$ solution, add 0.2mol/L NaOH solution dropwise until a colour change is observed.
6. To the same solution, now add 0.2mol/L solution dropwise until another colour change occurs. Record your observations.

Part B: The equilibrium between $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{CoCl}_4]^{2-}$.

1. Using the forceps, place a piece of the blue CoCl_2 paper into a 100mL beaker or Petri dish.
2. Add water drop-wise until a definite colour change is noted. Record your observations.
3. To the same piece of paper, add 1mol/L HCl dropwise until a definite colour change is noted. Record your observations.

Safety note:

CoCl_2 is toxic and a possible carcinogen. The indicator papers should be handled with forceps or gloves and hands washed after use.

Q: By referring to the collision theory, account for the observed colour change that occurred when HCl was added to the K_2CrO_4 solution.

Increasing $[\text{H}^+]$ by adding HCl decreases the distance between reactant particles, increasing frequency of successful collisions on the LHS, increasing the rate of the forward reaction relative to the reverse reaction, and the equilibrium shifts right until a new equilibrium is established. Thus, the solution turns more orange.

Q: When NaOH solution was added to a solution containing HCl , $[\text{H}^+]$ ions was reduced. By referring to the collision theory, account for the observed colour change that occurred when NaOH solution was added to the acidified K_2CrO_4 solution.

Decreasing $[\text{H}^+]$ by adding OH^- ions increases the distance between reactant particles, decreasing frequency of successful collisions on the LHS, decreasing the rate of the forward reaction relative to the reverse reaction, and the equilibrium shifts left until a new equilibrium is established. Thus, the solution turns more yellow.

Q: By referring to collisions theory, account for the observed colour change that occurred when NaOH solution and then the HCl solution were added to the $\text{K}_2\text{Cr}_2\text{O}_7$ solution.

Decreasing $[\text{H}^+]$ by adding OH^- ions increases the distance between reactant particles, decreasing frequency of successful collisions on the LHS, decreasing the rate of the forward reaction relative to the reverse reaction, and the equilibrium shifts left until a new equilibrium is established. Thus, the solution turns more yellow.

Q: Explain, with reference to collision theory, the colour changes that occurred when water and then HCl solution were added to the CoCl_2 paper.

Increasing $[\text{Cl}^-]$ by increasing $[\text{HCl}]$ decreases the distance between reactant particles, increasing frequency of successful collisions on the LHS, increasing the rate of the forward reaction relative to the reverse reaction, and the equilibrium shifts right until a new equilibrium is established. Thus, the solution turns more blue.

Q: When the equilibrium between $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{CoCl}_4]^{2-}$ ions is heated, it changes from pink to blue. Predict whether the reaction is exothermic or endothermic. Explain.

According to LCP, the system will act to partially counteract the imposed change. As such, increasing temperature will favour the endothermic reaction which decreases temperature. Since $[\text{CoCl}_4]^{2-}$ increases, the forward reaction is favoured, and so the forward reaction is endothermic.

Experiment 12: Preparation of Standard Na_2CO_3 Solution

Equipment:

- Balance.
- Volumetric flask (500mL).
- Oven.
- Desiccator.
- Beaker (250mL).
- Wash bottle.
- Storage bottle.
- Distilled water.
- Anhydrous Na_2CO_3 (4g).
- Pasteur pipette.

Procedure:

1. Calculate the mass of anhydrous Na_2CO_3 required to make up 500mL of 0.05mol/L solution.

2. Place a little more than the required amount in an oven at 270°C for 30 minutes to remove any water. After drying, place the anhydrous Na₂CO₃ in a desiccator to cool.
3. Accurately weigh out into a 250mL beaker the mass of Na₂CO₃ calculated. You don't need to make the mass exactly equal to that calculated but make sure you remember the mass so that the exact concentration can be calculated.
4. Dissolve the solid in around 100mL of distilled water.
5. Transfer the solution into a 500mL volumetric flask. Rinse the beaker several times with around 20mL portions of distilled water, adding each washing to the volumetric flask.
6. Make up the solution to precisely 500mL with distilled water, adding the last few millimetres dropwise from a Pasteur pipette.
7. Place the stopper in the volumetric flask and mix the solution thoroughly by repeatedly inverting the flask.
8. Transfer the solution to a clean storage bottle which should be first rinsed with a little of the Na₂CO₃ solution. Label the storage bottle with the type of solution.

Q: Calculate the precise solution concentrated in mol/L.

$$n(\text{Na}_2\text{CO}_3) = \frac{4}{M(\text{Na}_2\text{CO}_3)} = 0.0377\text{mol}$$

$$[\text{Na}_2\text{CO}_3] = \frac{0.0377}{0.5} = 0.0755\text{mol/L}$$

Q: The concentration of the solution you've made should be accurately known. List some possible ways that minor inaccuracies could occur in determining the concentration.

Rinsing the volumetric flask with water instead of the Na₂CO₃ solution.

Not completely transferring the Na₂CO₃ solution.

Experiment 13: Preparation and Standardisation of HCl

Equipment:

- Concentrated HCl (6mL).
- Graduated cylinder (10mL).
- Volumetric flask (500mL).
- Storage bottle.
- Beakers.
- Burette and stand.
- Funnel.
- Pipette.
- Pipette filler.
- Conical flask.
- Standard Na_2CO_3 solution.
- Methyl orange or bromophenol blue.
- Distilled water.
- Pasteur pipette.

Part A: Making the approximately 0.1mol/L HCl.

1. Calculate the volume of concentrated HCl that would be required to prepare 500mL of 0.1mol/L solution.
2. Measure out this volume of HCl in a graduated cylinder and transfer to a 500mL volumetric flask that's around half-filled with distilled water. Make the solution up to the mark using distilled water added dropwise from a Pasteur pipette.
3. Place the stopper in the volumetric flask and mix the solution thoroughly by repeatedly inverting the flask.
4. Transfer the approximately 0.1mol/L HCl solution to a clean storage bottle that has been rinsed with a little bit of the HCl solution.
5. Label the bottle with "0.1mol/L HCl".

Safety note:

- Concentrated HCl is very corrosive and must be handled with extreme care. You must wear eye protection.
- If any concentrated HCl gets in contact with your skin, immediately wash it off with copious quantities of water.

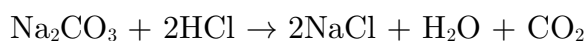
Part B: Standardisation of the HCl acid solution.

1. Place around 100mL of the standard Na_2CO_3 solution into a clean beaker. If the beaker is wet, rinse with a little bit of the Na_2CO_3 solution first.
2. Rinse a clean 20mL pipette with some of the Na_2CO_3 solution. Pipette 20mL of the Na_2CO_3 solution into a 250mL conical flask. Add 2-3 drops of your chosen indicator to the flask.
3. Place around 100mL of the newly made HCl solution into a clean beaker. Again, if necessary, rinse the beaker with a bit of the HCl solution first.
4. Rinse a clean burette with some of the HCl solution and then fill the burette with the solution.
5. Note and record the level of acid in the burette. Obtain a rough estimate of the titration volume by running acid quickly from the burette while constantly swirling the liquid in the conical flask. Stop delivery of the acid as soon as a permanent colour change is obtained. Note and record the acid level in the burette and determine the approximate volume of acid required.
6. Record your results in a table.
7. Prepare another conical flask containing a 20mL aliquot of the Na_2CO_3 solution and 2-3 drops of indicator. This time, add the acid quickly from the burette with constant swirling of the flask, until the volume added is within 2-3mL of the approximate volume required. Rinse the inside of the conical flask with a jet of distilled water from a wash bottle to return any splashed solution to the bulk solution. Continue adding acid dropwise and with constant swirling until the addition of one drop is sufficient to produce a permanent

colour change. Note and record the level of acid in the burette at the endpoint.

8. Repeat the accurate titration with further 20mL aliquots of Na_2CO_3 solution until consistent titration volumes are obtained. These should be within 0.3mL of each other.

Q: Using the equation for the reaction, calculate the number of moles of Na_2CO_3 used in each titration.



Q: Distinguish between equivalence point of a reaction and endpoint of a titration.

Equivalence point is when the number of moles of acid is equal to the number of moles of base. Endpoint is when the indicator changes colour.

Q: Suppose that phenolphthalein, whose colour change is in the vicinity of pH 9, had been used instead of one of the indicators recommended. Would:

[a] The volume of acid required for the titration be more or less than that obtained in this experiment?

Less because less acid would be required to make the solution a pH of 9.

[b] The calculated concentration of the HCl solution be higher or lower than the result obtained in this experiment?

Less because decreasing the average titre volume of HCl would mean a lower calculated number of moles of Na_2CO_3 and hence a lower concentration.

Q: What are possible sources of error in this experiment?

- Irregular changes in temperature and pressure in the surroundings.
- Endpoint determination – a visual endpoint is always slightly beyond the equivalence point due to needing to see the colour change.
- Parallax error.
- Inherent error in the burette.
- Inherent error in the conical flask.
- Inherent error in the pipette.

Experiment 19: Oxidation and Reduction Reactions Involving Metals

Equipment:

- 3 test tubes.
- Samples of copper, lead, magnesium and zinc (4 samples each).
- 2mol/L HCl (10mL).
- Taper and matches.
- 10mL 0.1mol/L solutions of:
 - $\text{Cu}(\text{NO}_3)_2$
 - $\text{Pb}(\text{NO}_3)_2$
 - $\text{Mg}(\text{NO}_3)_2$
 - $\text{Zn}(\text{NO}_3)_2$

Part A: Action of dilute HCl on various metals.

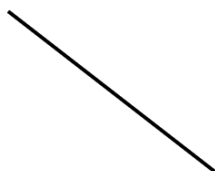
1. Place a piece of zinc metal in a test tube and add around 2mL of dilute HCl. Observe what happens.
2. Collect any gas that has evolved in another test tube by inverting it over the first test tube. Apply a lighted taper to the mouth of the inverted test tube and identify the gas present from the result of this test.
3. Repeat this procedure for magnesium, lead and copper, noting any difference in behaviour from the zinc.

Part B: Metal-metal ion displacement reactions.

1. Place a piece of zinc in each of the 3 test tubes and add sufficient $\text{Mg}(\text{NO}_3)_2$ solution to the first test tube, $\text{Pb}(\text{NO}_3)_2$ solution to the second test tube and $\text{Cu}(\text{NO}_3)_2$ to the third test tube to cover the zinc. Observe and note any reactions that occur.
2. Repeat this procedure for the magnesium, covering the metal with $\text{Zn}(\text{NO}_3)_2$, $\text{Pb}(\text{NO}_3)_2$ and $\text{Cu}(\text{NO}_3)_2$ solutions respectively.
3. Repeat this procedure for the lead, covering the metal with $\text{Zn}(\text{NO}_3)_2$, $\text{Mg}(\text{NO}_3)_2$ and $\text{Cu}(\text{NO}_3)_2$ solutions respectively.
4. Repeat this procedure for the copper, covering the metal with $\text{Mg}(\text{NO}_3)_2$, $\text{Pb}(\text{NO}_3)_2$ and $\text{Zn}(\text{NO}_3)_2$ solutions respectively.

Q: Write the oxidation and reduction half-equations and the overall redox equation for any reactions that took place.

Metal ion



Solid metal

Q: For each reaction you observed, identify the oxidising and reducing agents.

Q: List the 4 metals observed in this experiment in order of decreasing strength as reducing agents.

Equipment:

- 6 test tubes and stoppers.
- 5mL 0.5mol/L solutions of:
 - KBr.
 - KCl.
 - KI.
- 5mL chlorine water.
- 5mL bromine water.
- 5mL iodine water.
- 10mL dichloromethane.

Part A: Colour of halogens in dichloromethane.

1. Prepare a table to record the colours of the halogens and halide ions in water and the halogens in dichloromethane.
2. Record the colour of the bromine water which contains the halogen molecule, Br₂. Record the colour of the KBr solution which contains Br⁻ ions.
3. Place around 2mL bromine water in a test tube. Add 1mL dichloromethane, stopper and shake vigorously for a few seconds. Allow liquid layers to separate. Record the colour of the dichloromethane layer (the bottom layer) which contains the halogen molecule, Br₂.
4. Repeat procedures 2 and 3 for chlorine water and iodine water.

Safety note:

- The halogen solutions are poisonous and must be handled with care.
- Don't breathe the vapours given off from these solutions.
- If the solutions come into contact with your skin, immediately wash the affected area with plenty of water.
- Dichloromethane is poisonous and should be handled with care.

- Don't breathe in dichloromethane vapour or allow it to come into contact with your skin.
- Carry out the experiment in the fume hood.

Part B: Halogen displacement reactions.

1. Add around 2mL of chlorine water to 2mL of KBr solution in a test tube and shake. Observe any changes that occur.
2. Add around 1mL of dichloromethane to the same test tube. Shake and allow the dichloromethane layer to settle, and note its colour. From the colour of the dichloromethane, can you identify the halogen it has dissolved.
3. Repeat this procedure for solutions of the other halogens and halide ions, using the following combinations.

Chlorine water + KI solution.

Bromine water + KCl solution.

Bromine water + KI solution.

Iodine water + KCl solution.

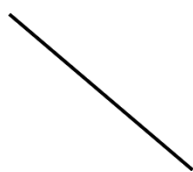
Iodine water + KBr solution.

In each case, add some dichloromethane and shake and identify any newly formed halogen that might be present.

Mixture:	Colour:	Halogen dissolved:
Chlorine water + KI solution		
Bromine water + KCl solution		
Bromine water + KI solution		
Iodine water + KCl solution		
Iodine water + KBr solution		

Q: Write the oxidation and reduction half-equations and the overall redox equation for any reactions that took place.

Halogen



Halide ion

Q: Which of the halogens is the strongest oxidising agent?

Q: List the halide ions in order of increasing strength as reducing agents.