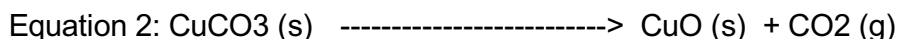
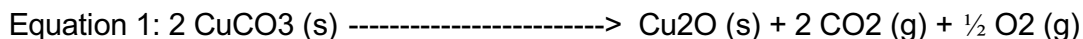


## Chemistry Coursework Skill P: Decomposition of copper carbonate

This is an experiment to determine the thermal decomposition of copper carbonate. As seen in equation 1 and 2 Cu<sub>2</sub>O and CuO are possible outcomes.



By performing a quantitative analysis of the gas volume the two equations can be distinguished.

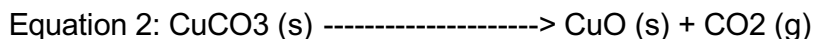
### Background theory

Copper carbonate is found in nature as the mineral malachite and is a green powder. When heated it decomposes to give black copper oxide (CuO) and carbon dioxide. The melting point is 260 °C. Copper has two different oxides, CuO and Cu<sub>2</sub>O. CuO transfers at 900 °C into Cu<sub>2</sub>O with loss of oxygen (Holleman - Wiberg). Thus for this experiment I expect equation 2 to be correct.

### Calculating the correct mass of copper carbonate

To receive an adequate amount of gas volume (maximum volume of burette: 50 cm<sup>3</sup>) I first have to calculate the necessary amount of CuCO<sub>3</sub>:

With a molar ratio of 1:1 I use Equation two to calculate the necessary mass of copper carbonate



'1 mole of copper carbonate will decompose to 1 mole of CuO and 1 mole of CO<sub>2</sub>'

'1 mole of gas occupies 24360 cm<sup>3</sup> at 20°C, 100kPa'.

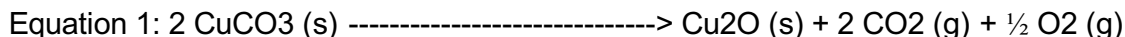
I aim at 30 cm<sup>3</sup> volume of gas, so I can ensure all gas is measured.

Thus, if 1 mole of copper carbonate produces 24360 cm<sup>3</sup> of CO<sub>2</sub>, to produce 30 cm<sup>3</sup> we need:

- number of moles =  $\frac{\text{volume cm}^3}{24360 \text{ cm}^3}$   
=  $\frac{30 \text{ cm}^3}{24360 \text{ cm}^3}$   
= 0.0012315271 moles of gas
- Mr of CuCO<sub>3</sub> = 63.5+12+(3\*16)= 123.5
- Mass= moles\* Mr = 0.0012315271 moles\* 123.5 = 0.1520935969 g = 0.15 g ( 2 d.p.)  
of CuCO<sub>3</sub>

With equation 2 0,15g of  $\text{CuCO}_3$  will result in  $30 \text{ cm}^3$  gas.

I now calculate the expected volume of gas given equation one is correct:



- The molar ratio in equation one is 2:2.5
  - number of moles = mass/Mr  
 $= 0.1520935969 \text{ g} / 123.5 = 0.0012315271$  moles of copper carbonate
  - $0.0012315271 \text{ moles} * 1.25 = 0.0015394089$  moles of gas
- > to find moles of gas I multiply mole s of  $\text{CuCO}_3$  by 1.25, as  $2.5/2 = 1.25$ .
- $0.0015394089 \text{ moles} * 24360 \text{ cm}^3 = 37.5 \text{ cm}^3$

At  $20^\circ\text{C}$ , 100kPa the two possible results are 30 and  $37.5 \text{ cm}^3$  respectively.

#### Fair test/ Sources of Error

- copper carbonate has to be weighed out exactly to two decimal place s
- bung has to be inserted tightly into test tube before the reaction occurs
- ensure that delivery tube is underneath burette and that burette is vertical
- do not stop reaction as long as bubbles are created
- wait circa 10 min (gas to reach room temperature) a fter ending of reaction to measure gas evolved
- ensure that there are no air bubbles in the burette
- perform experiment accordingly without  $\text{CuCO}_3$  and measure collected gas.  
This volume is to be deducted from the actual results (temperature effect)
- ensure that all copper carbonate has been fully decomposed (colour change)
- precise reading for temperature and air pressure

#### Variables

controlled variables:

- the pressure of the room, to measure with a barometer for the final calculation.
- the temperature. An increase in temperature may cause an increase in the volume of gas given out by the expansion of the air.
- degree of decomposition of copper carbonate. The green copper carbonate has to turn black completely and has to be heated evenly(heating time).
- 

Dependent variable:

- amount of gas.
- 

Independent variable

- copper carbonate - fixed

#### Risk assessment

- copper carbonate is toxic and it can irritate lungs and eyes
- copper(1)oxide is harmful if swallowed and can irritate lungs and eyes
- carbon dioxide has a low hazard symbol. It is asphyxiate and lowers the oxygen content of the air.

- use the Bunsen burner careful and turn on safety flame when not in use
- wear goggles and lab coat all the time.
- when something is spoiled it should be rinsed up immediately
- wear your hair looped up
- apparatus may be hot

### Apparatus list

- 1 Bunsen burner for heating
- 1 test tube, good for small weighted sample and, only little air enclosed
- 1 delivery tube to deliver gas to burette
- 1 pair of goggles for safety
- 1 lab coat for safety
- 2 clamp stands for holding test tube and burette
- 1 50 cm<sup>3</sup> burette, accurate to 0.05 cm<sup>3</sup>, used to collect the evolved gas
- 1 heat mat underneath Bunsen burner
- 1 bung to close test tube, leak proofed
- 1 spatula to deliver CuCO<sub>3</sub>
- 1 top pan balance, accurate to 0.005g, is used to weigh copper carbonate
- 0.15g copper carbonate
- 150 cm<sup>3</sup> distilled water
- 1 100 cm<sup>3</sup> beaker, container to keep end of burette under water
- 1 sheet of paper
- 1 Thermometer for correct measurement of temperature
- barometer (not available)
- 

### Sketch

### Method

- 1 wear goggles and lab coat. Pin hair up
- 2 collect all apparatus in front of you
- 3 set up apparatus as shown in the sketch above
- 4 put a piece of paper on the top pan balance and reset the scale
- 5 put copper carbonate with a spatula onto the piece of paper and weigh exactly 0.15g

- 6 transfer  $\text{CuCO}_3$  into middle of test tube and insert the bung tightly (no air gaps)
- 7 measure temperature
- 8 begin heating powder gently with Bunsen burner
- 9 gas is led into vertical burette
- 10 when no more gas bubbles are formed, lift end of delivery tube out of water and stop heating
- 11 after 10 minutes read and note exactly volume of gas (reading minus 50  $\text{cm}^3$ ) (2 d.p.)
- 12 repeat experiment three times
- 13 work out average volume of gas evolved

The volume of gas is temperature and pressure dependent. Therefore you have to use a barometer (not available at school) and a thermometer to calculate the correct volume. I reduce the temperature effect of the heated gas by waiting 10 minutes before measuring it.

### General gas equation

$$P \cdot V = n \cdot R \cdot T$$

P= Pressure

V= Volume

n=moles of gas

R= gas constant =  $8.314 \text{ J K}^{-1} \text{ mol}^{-1}$

T= temperature

Example calculation:

$$P \cdot V = n \cdot R \cdot T$$

$$V = (n \cdot R \cdot T) / P$$

$$= (1 \text{ mol} \cdot 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \cdot 293 \text{ K}) / 100 \text{ kPa}$$

$$= 24.36 \text{ dm}^3 \text{ at } 20^\circ \text{C and } 100 \text{ kPa pressure}$$

After calculating an average volume of gas produced you can compare this with the theoretical results of  $30 \text{ cm}^3$  for equation two and  $37.5 \text{ cm}^3$  for equation one. As mentioned above I predict equation 2 to be correct, thus  $30 \text{ cm}^3$ , or closest to this, of gas should be measured. Due to the small amount of reactants already small deviations in measurement of weight, temperature and pressure could, when added up, show a certain margin of error. Thus more accurate tools (e.g. analytical balance), increased mass of reactants and standardized heating will improve accuracy of results. Alternatively use gravimetric measurement of  $\text{CuCO}_3$  and  $\text{CuO}$  respectively.

### Bibliography

- Hazard references from the school laboratory
- Personal class notes
- OCR chemistry, page 1-20
- [www.wikipedia.org/wiki/copper%28](http://www.wikipedia.org/wiki/copper%28)
- Advanced Chemistry by Michael Clugston and Rosalind Flemming, page 114-122
- Holleman- Wiberg, Lehrbuch der Anorganischen Chemie, page 784 -796

- Hans Rudolf Christen, Chemie, page 147 -150