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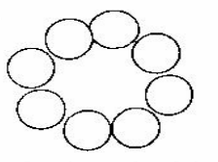
INORGANIC CHEMISTRY

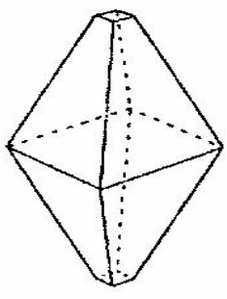
A. Sulphur and sulphur compounds:

1. Sulphur(S)(yellow colour),occurring in more than one form, is called **allotropism**, for example:

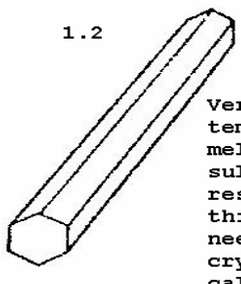
1.1



S8-ring rhombic sulphur

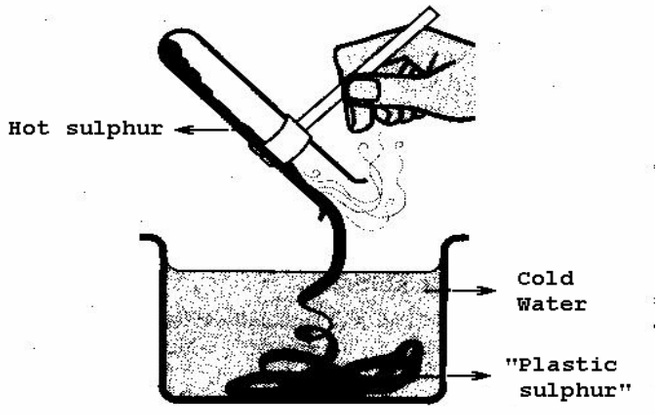


1.2



Very low temperature melting of sulphur results in this long needle-like crystals, called monoclinic sulphur.

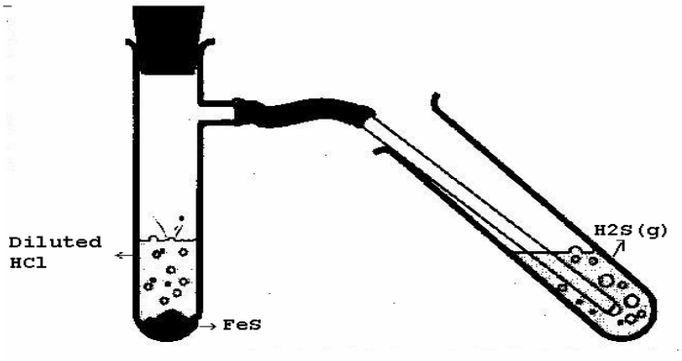
1.3



2. The sulphur compounds:

2.1 Hydrogen sulphide (H_2S):

2.1.1 Preparation: Pour diluted hydrochloric acid (HCl)on a metal sulphide, say, FeS (Black coloured pieces), as follows:



Reaction: $FeS + HCl(dil) \rightarrow H_2S(g) + FeCl_2$ (Unbalanced)

$FeS + 2HCl \rightarrow H_2S + FeCl_2$ (Balanced by inspection)

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2.1.2 Properties of H_2S:
Denser than air, "rotten egg" smell and poisonous.

2.1.3 Test for H_2S:
Blackening of a filter paper dipped in lead acetate solution:

$$Pb(CH_3COO)_2 - H_2S(g) \rightarrow \underset{\text{Black}}{PbS} + 2CH_3COOH$$

2.1.4 Other precipitation reactions with H_2S

a) $CuSO_4 + H_2S \rightarrow \underset{\text{Black}}{CuS} + H_2SO_4$

b) $Pb(NO_3)_2 + H_2S \rightarrow \underset{\text{Black}}{PbS} + 2HNO_3$

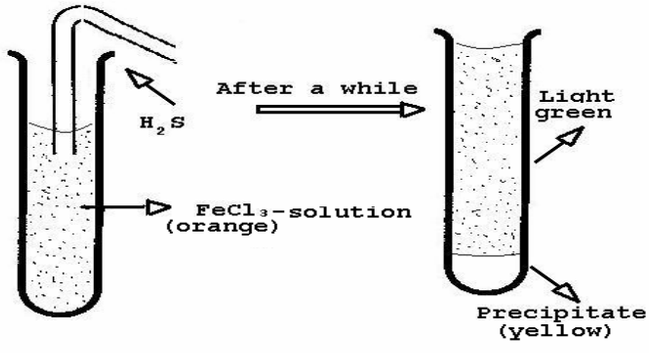
c) $ZnCl_2 + H_2S \rightarrow \underset{\text{White}}{ZnS} + 2HCl$ (Test for Zn^{2+}-ions)

d) $2Ag + H_2S \rightarrow Ag_2S + H_2$ (Tarnish of silverware)

2.1.5 H_2S as reducing agent in oxidation reactions

a) H_2S easily forms ions in solutions as follows:
$H_2S \rightarrow 2H^+ + H_2S \rightarrow S^{2-}$
or (from Table: $H_2S \rightarrow S + 2H^+ + 2e^-$)

Example:



1. First give preparation of H_2S, then test and then smell (see p.1)and other properties.
2. Give the oxidation half-reaction (use table)
   $H_2S \rightarrow S + 2H^+ + 2e^-$
   (Remember: Oxidation is "giving of" e^-)
3. Give the reduction half-reaction in the test tube.
   $Fe^{3+} + e^- \rightarrow Fe^{2+}$
   (Remember: Reduction is the "taking up" of e^-)
4. Write down the balanced ionic reaction in the test tube:
   (Remember: Get e^- equal in number and add up both equations)
   $H_2S \rightarrow S + 2H^+ + 2e^-$ (Oxidation)
   $2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+}$ (Reduction)
   $2Fe^{3+} + H_2S \rightarrow 2Fe^{2+} + S + 2H^+$
   Orange → Green

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5. Precipitate: Sulphur(S)(See the balanced ionic reaction).

6. Ion responsible for green colour: Fe^{2+}

b. Other reactions (with H_2S as reducing agent):

b.i) Potassium permanganate solution:

$H_2S \rightarrow S + 2H^+ + 2e^-$ (See table)

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (See table)

Now balance electrons (and add up):

$5H_2S \rightarrow 5S + 10H^+ + 10e^-$

$2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$

$2MnO_4^- + 5H_2S + 16H^+ \rightarrow + 5S + 10H^+ + 2Mn^{2+} + 8H_2O$

(purple) (Precipitate) (Colourless)

b.2) Potassium dichromate solution:

$H_2S \rightarrow S + 2H^+ + 2e^-$ (Oxidation)

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ (Reduction)

Now balance electrons:

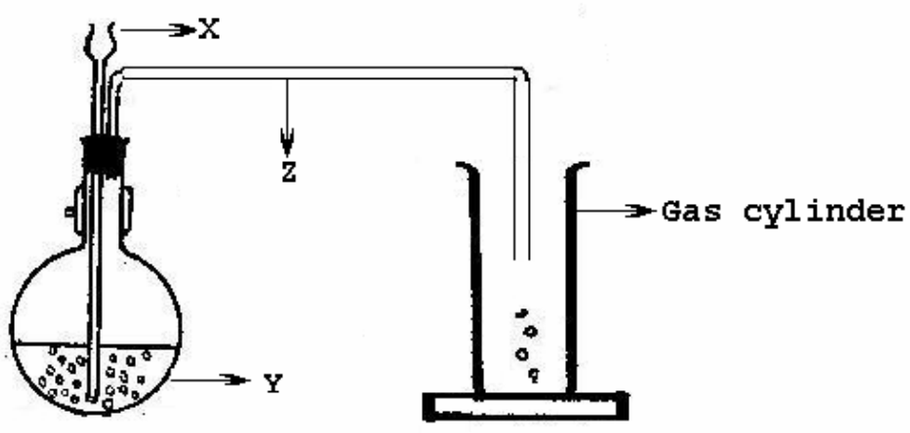
$3H_2S \rightarrow 3S + 6H^+ + 6e^-$

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

$3H_2S + Cr_2O_7^{2-} + 14H^+ \rightarrow 3S + 2Cr^{3+} + 6H^+ + 7H_2O$

(Orange) (Green)

Summarizing question (about H_2S)



1. What is X, Y and Z?

Answer:

X is HCl, Y is FeS and Z is H_2S.

2. Give the balanced reaction.

Answer:

$FeS + 2HCl \rightarrow FeCl_2 + H_2S$

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3. Give two properties of H_2S.

Answer:

Poisoness, denser than air, "bad smell".

4. Is the way of collecting the gas correct?

Answer:

Yes, it is denser than air and it is called "the upward displacement of air by the denser gas H_2S".

2.2 Sulphur dioxide (SO_2):

2.2.1 Sulphur(s) burning in oxygen forms SO_2: $S+O_2 \Rightarrow SO_2$

2.2.2 Copper (Cu) and concentrated sulphuric acid (H_2SO_4):

$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + 2H_2O + SO_2$$

2.2.3 In the laboratory: Diluted hydrochloric acid (HCl) with sodium sulphite (Na_2SO_3) or sodium hydrogen sulphite ($NaHSO_3$):

$$Na_2SO_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2$$

$$NaHSO_3 + HCl \rightarrow NaCl + H_2O + SO_2$$

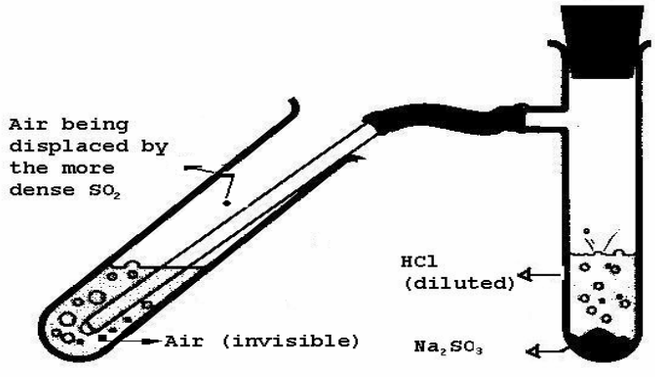
2.2.4 Test: $KMnO_4$ (purple) - solution discolours (loses colour) with SO_2.

2.2.5 Properties: Poisonous and dissolves easily in water (H_2O):

$$SO_2(g) + H_2O(l) \rightarrow H_2O_3$$

(Sulphurous acid turns blue lithmus red).

Heavier (more dense) than air → that the gas is collected by means of the downward displacement of air as follows:



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2.2.6 Chemical reactions of SO_2:

I: As reducing agent (in oxidation reactions):

Bubble SO_2 through potassium permanganate and potassium dichromate solutions (and note how they discolour).

a) $SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$ (Oxidation)

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ (Reduction)

$5SO_2 + 10H_2O + 2MnO_4^- + 16H^+ \rightarrow 5SO_4^{2-} + 20H^+ + 2Mn^{2+} + 8H_2O$

Simplify:

$5SO_2 + 2H_2O + MnO_4^- \rightarrow 5SO_4^{2-} + 4H^+ + 2Mn^{2+}$

(purple colour; colourless)

b) $SO_2 + 2H_2O + \rightarrow SO_4^{2-} + 4H^+ + 2e^-$ (Oxidation)

$Cr_2O_7^- + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ (Reduction

$3SO_2 + 6H_2O + Cr_2O_7^- + 14H^+ \rightarrow 3SO_4^{2-} + 12H^+ + 2Cr^{3+} + 7H_2O$

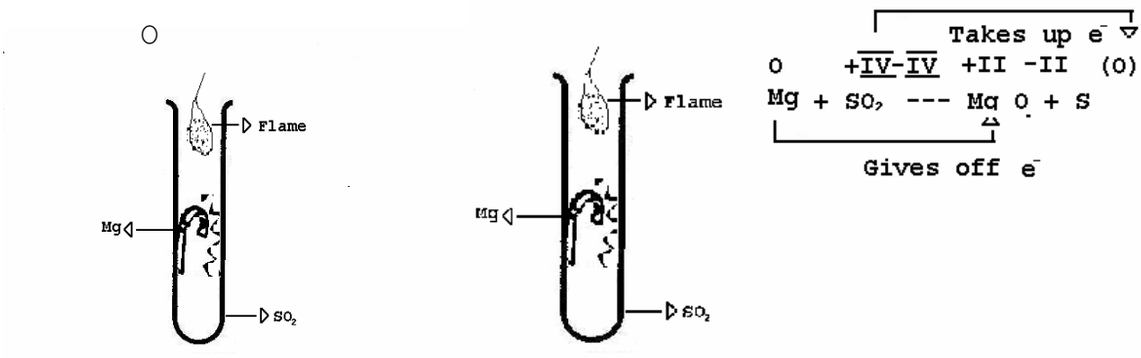
Simplify:

$3SO_2 + Cr_2O_7^- + 2H^+ \rightarrow 3SO_4^{2-} + 12H^+ + 2Cr^{3+} + H_2O$

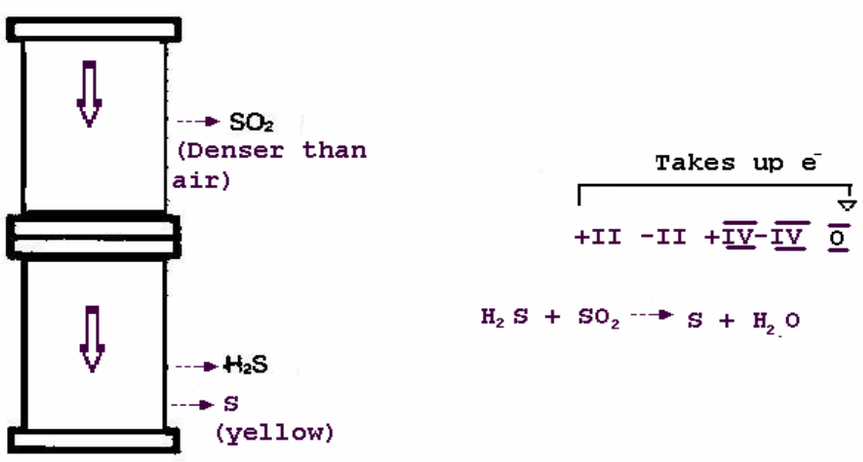
(orange; yellow)

II: Oxidizing agent (in reduction reactions):

a.



b.



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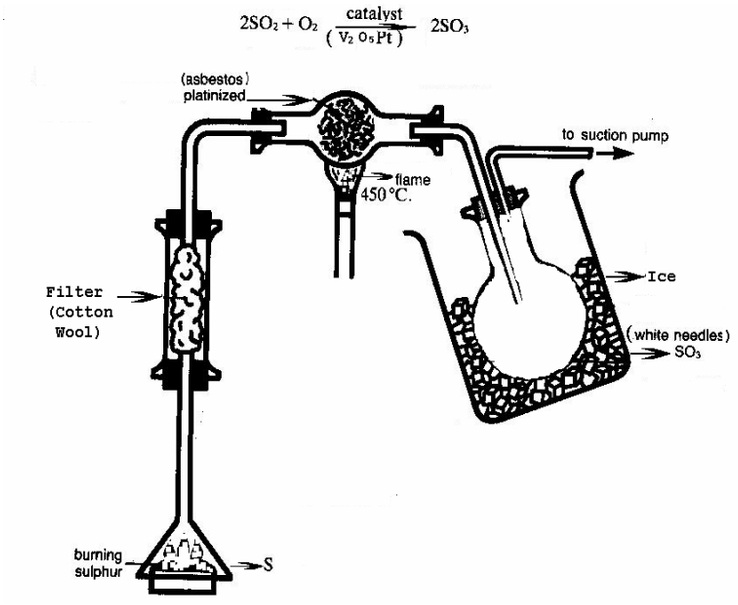
2.3 Sulphur trioxide (SO_3):



Properties of SO_3:

Extremely hygroscopic (removes H_2O from the air):

$SO_3 + H_2O \rightarrow H_2SO_4$ (contact process).

2.4 Sulphuric acid (H_2SO_4):

1. Preparation: (The contact process)

$S + O_2 \rightarrow SO_2$ (gas)

$2SO_2 + O_2 \xrightarrow[\underline{\text{or}}\ Pt]{\text{catalist } V_2O_5} SO_3$ (needles)

$SO_3 + H_2SO_4(c) \rightarrow H_2S_2O_7$ (Oleum or pyro-sulphuric acid)

$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$ (Sulphuric acid)

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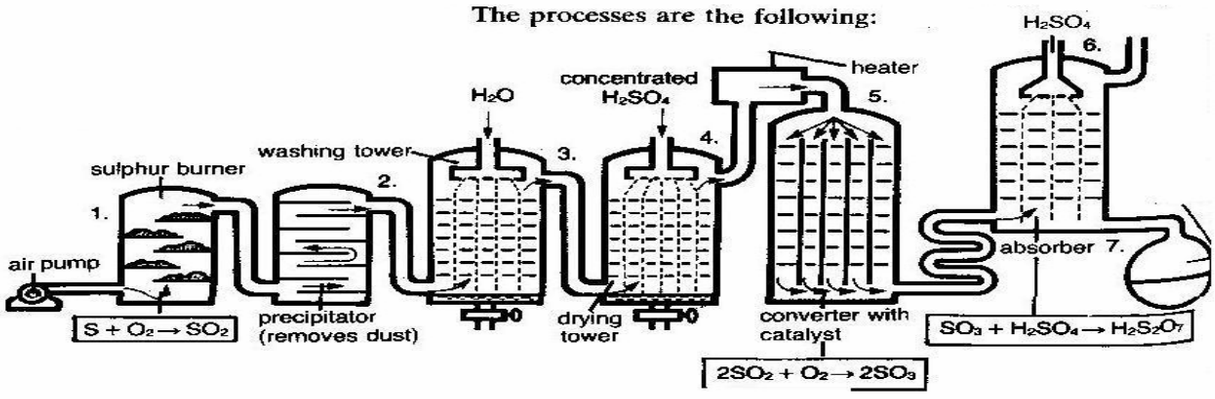
Preparation of sulphuric acid using the contact process

The processes are the following:



ii. Properties (of H_2SO_4)

a) Dehydrating property:

$$C_{12}H_{22}O_{11} \xrightarrow{H_2SO_4\ (c)} 12\ C + 11H_2O$$

(Sugar)

$$CuS\ O_4.5H_2O \xrightarrow{H_2SO_4\ (c)} CuSO_4 + 5H_2O$$

Blue White

Remember: (Test for H_2O)

$$CuSO_4.5H_2O \xrightarrow{\Delta H} CuSO_4 + 5H_2O$$

Blue **(Heat)** **White**

b) Oxidizing agent:

$$SO_4^{2-} + 4H^+ + 2e^- \xrightarrow{\Delta H} SO_2 + 2H_2O$$

$$Cu \rightarrow Cu^{2+} + 2e^-$$

$$SO_4^{2-} + 4H^+ + Cu \rightarrow Cu^{2+} + SO_2 + 2H_2O$$

c) Preparations of acids from their corresponding salts:

$$NaCl + H_2SO_4\ (c) \rightarrow Na_2SO_4 + 2HCl$$

$$2KNO_3 + H_2SO_4\ (c) \rightarrow K_2SO_4 + 2HNO_3$$

(Examples of double exchange of ions).

iii. Uses of H_2SO_4:

a) In car batteries; b) Production of fertilizers; c) Cleaning of metals; d) Recovery of uranium.

iv. Uses of its salts:

a) Plaster of Paris ($CaSO_4$)

b) Medicines; ($MgSO_4$) and Na_2SO_4).

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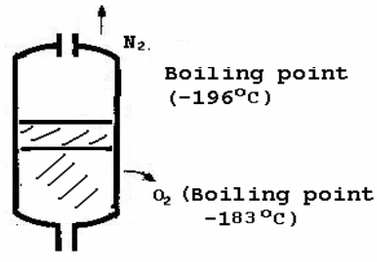
v. Test for the sulphates:

Add a few drops of $BaCl_2$ or $CaCl_2$ solutions to the sulphates. A white precipitate forms which will not dissolve in HCl.

3. The nitrogen compounds:

3.1 Nitrogen (N_2):

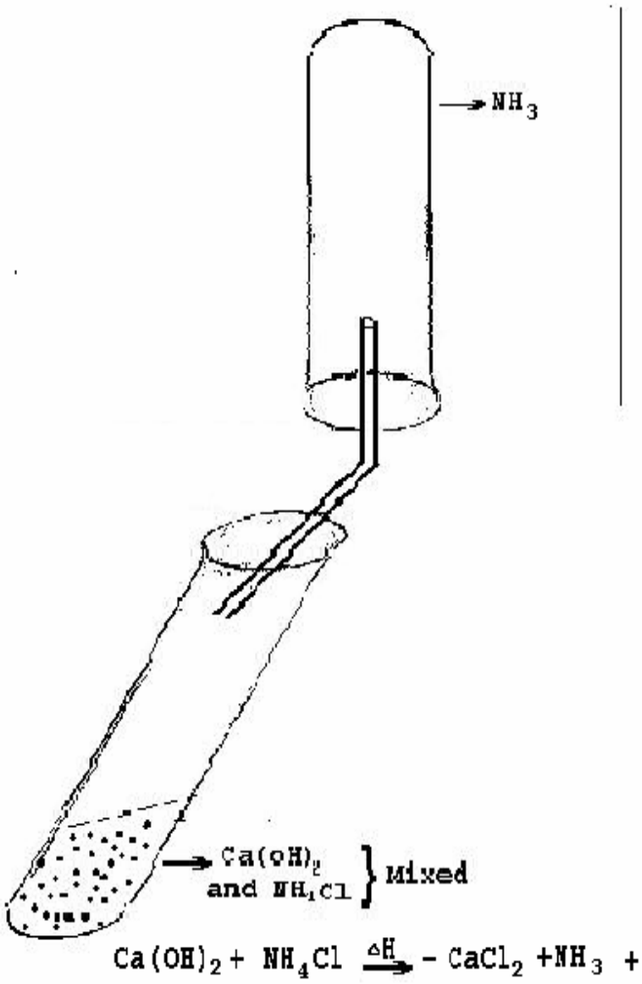
Liquefied air is fractionally distilled accordingly to the different boiling points, eg.

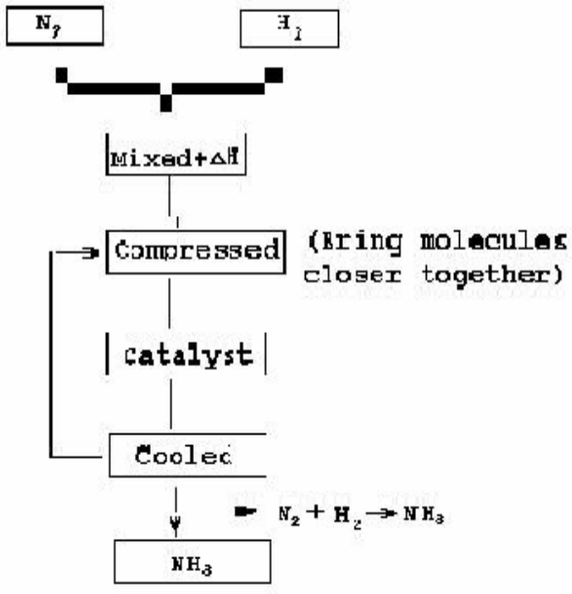


3.2 Ammonia (NH_3):

3.2.1 Preparation: (Heber process)

Laboratory: Industry:





$$Ca(OH)_2 + NH_4Cl \xrightarrow{\Delta H} -CaCl_2 + NH_3 + 2H_2O$$

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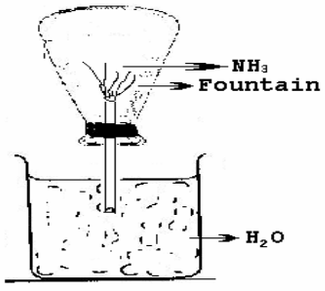
3.2.2 Properties (of NH_3):

i. Extremely soluble (dissolvable) in water:



$NH_3 + H_2O \rightarrow NH_4OH$

Alkali (Basic)

The "Fountain experiment".

ii. Forms a alkaline" or "basic" solution with H_2O (turns litmus blue) - test for NH_3.

3.2.3 Uses (of NH_3)

Cleaning agent, preparation of saltpetre, fertilizers, and Ammonium salts.

The ammonium salts:

i. Formation of the salts of ammonia (formed by strong acids): $NH_3 + HCl \rightarrow NH_4Cl$

- $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$
- $NH_3 + HNO_3 \rightarrow NH_4NO_3$

ii. Properties of the ammonium salts:

- Very soluble in water.
- Break up (decompose) to form $NH_3(g)$:

$$NH_4HCO_3 \xrightarrow{\Delta H} NH_3\uparrow + H_2O + CO_2$$

$$NH_4Cl \xrightarrow{\Delta H} NH_3\uparrow + HCl$$

$$NH_4NO_3 \xrightarrow{\Delta H} N_2O + 2H_2O$$

(Salt from strong oxidizing agent)

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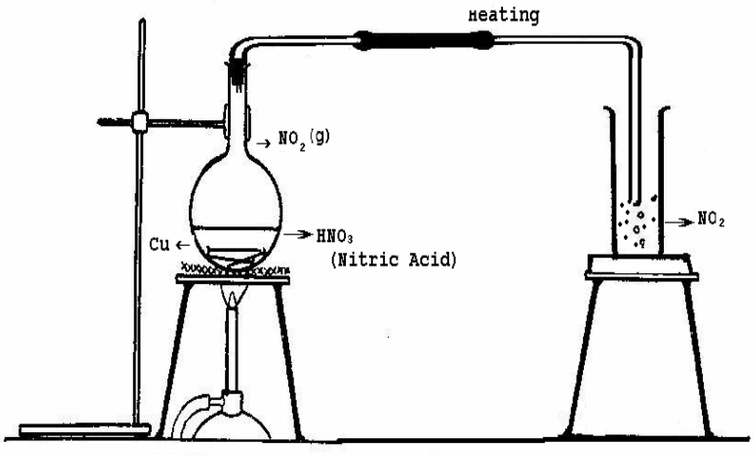
iii. Uses: Fertilizers and in soldering.

3.3 Nitrogen dioxide (NO_2): (Nitrogen-IV-oxide: NO_2)

3.3.1 Preparation:

$$2Pb(NO_3)_2 \xrightarrow{\Delta H} 2PbO + 4NO_2 + O_2$$
(Salt)



$$Cu + 4HNO_3^{(c)} \xrightarrow{\Delta H} Cu(NO_3)_2 + 2H_2O + 2NO_2$$

Using the table:

$Cu \longrightarrow Cu^{2+} + 2e^-$ (oxidation)

$NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$ (reduction)

$Cu + 2NO_3^- + 4H^+ \rightarrow Cu^{2+} + 2NO_2 + 2H_2O$ (redox)

3.3.2 Properties: (NO_2): Brown poisonous gas.

3.3.3 Equilibrium:

$$2NO_2 \rightleftharpoons N_2O_4 + \text{Energy}$$

(Nitrogen di-oxide) $\rightleftharpoons$ (Di-nitrogen tetroxide) (Exothermic)

Reddish Brown Light Brown to colourless

- Increase the pressure: To decrease the higher pressure, more molecules must be made less: 2 moles become 1 mole and the forward reaction speeds up, making more N_2O_4 (light brown).

- Decrease the pressure: To increase the lowered pressure, less molecules must be made more: 1 mole becomes 2 moles and therefore the reverse (backward) reaction speeds up, making more NO_2 (reddish brown).

- Increase the temperature: To decrease the temperature, heat (or energy) must be removed. The backward (reverse) reaction speeds up, removes the additional heat and forms more reactants, e.g. NO_2 (reddish brown).

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- Decrease the temperature: To increase the temperature, heat (or energy) must be formed. The forward reaction speeds up, forming heat (or energy) and in doing so also forming more products, e.g. N_2O_4 (Light brown to colourless).

3.4 Nitric acid (HNO_3):

1. Preparation (Laboratory):

$$2NaNO_3 + H_2SO_4\ (c) \rightarrow Na_2SO_4 + 2HNO_3$$

(Sodium Nitrate)

Preparation (Industry):

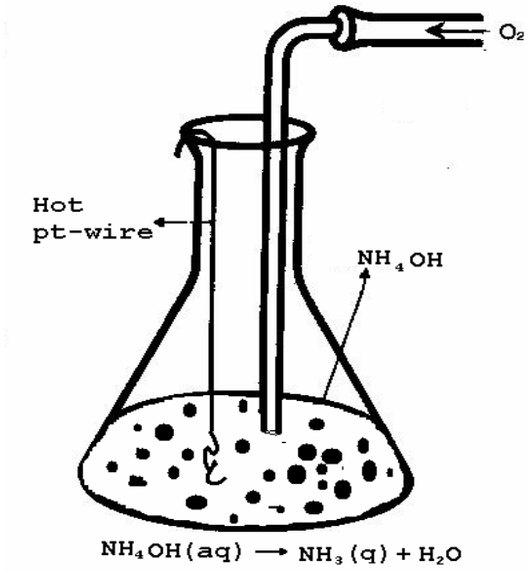
(The Ostwald process)

Step 1: $4NH_3 + 5O_2 \xrightarrow{\text{Catalyst}} 4NO + 6H_2O$

Step 2: $2NO + O_2 \longrightarrow 2NO_2$

Step 3: $4NO_2 + 2H_2O + O_2 \longrightarrow 4HNO_3$

The catalytic oxidation of NH_3:



When the $NH_3(q)$ forms, it is converted to NO immediately: See Step 1.

Now: $4NO + 3O_2 + 2H_2O \longrightarrow 4HNO_3$

The exothermic reaction takes place vigorously in the conical flask.

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ii. The decomposition of HNO_3 by heat (sunlight or flame):

$$4HNO_3 \xrightarrow{\Delta_H} 4NO_2 + 2H_2O + O_2$$

(Reddish Brown)

iii. Oxidizing action of nitric acid:

$$Cu \longrightarrow Cu^{2+} + 2e^- \text{ (Oxidation)}$$
$$NO_3^- + 2H^+ + 2e^- \longrightarrow NO_2(g) + H_2O$$
$$NO_3^- + 2H^+ + Cu \longrightarrow Cu^{2+} + NO_2 + H_2O$$

iv. Uses of nitric acid:

Making of explosives (like TNT)

and fertilizers (like $NH_4 \quad NO_3$).

v. The salts of nitric acid:

The salts of nitric acid (like that of HCl and H_2SO_4) are identified by the nitrate part of the salt, eg. KNO_3 (Potassium nitrate) and $NaNO_3$ (sodium nitrate).

vi. Heating the salts:

$$2NaNO_3 \xrightarrow{\Delta_H} 2NaNO_2 + O_2\uparrow(g)$$

Sodium nitrate — Sodium nitrite

3. Halogens and the halides:

3.1 Introduction:

The halogens (Group 7-elements) are fluorine(F), chlorine(Cl); bromine(Br) and iodine(I) - as well as astatine(At).

3.2 Properties:

i. Chlorine (Cl_2) - light yellowish gas.

Bromine (Br_2) - volatile reddish-brown liquid.

Iodine (I_2) - soft flaky crystals with a purplish-black colour.

ii. Poisonous and very reactive.

iii. Form the salts (called the halides), eg. NaCl, KCl, NaBr, NaI, KI.

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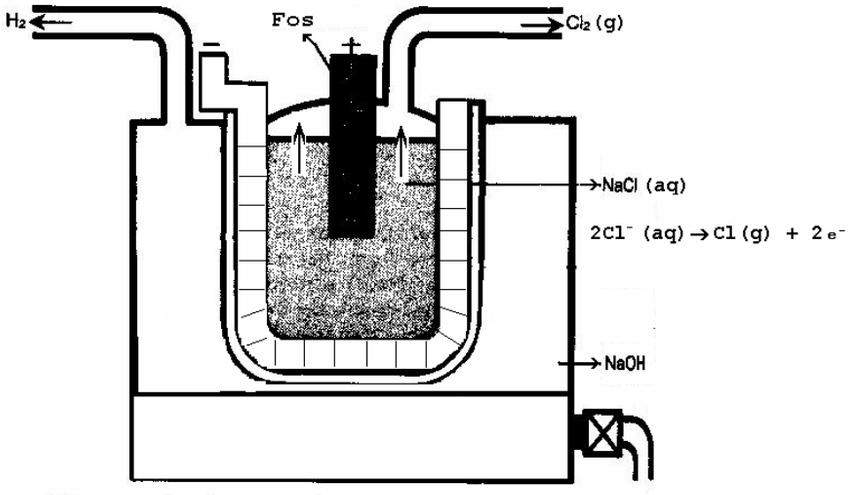
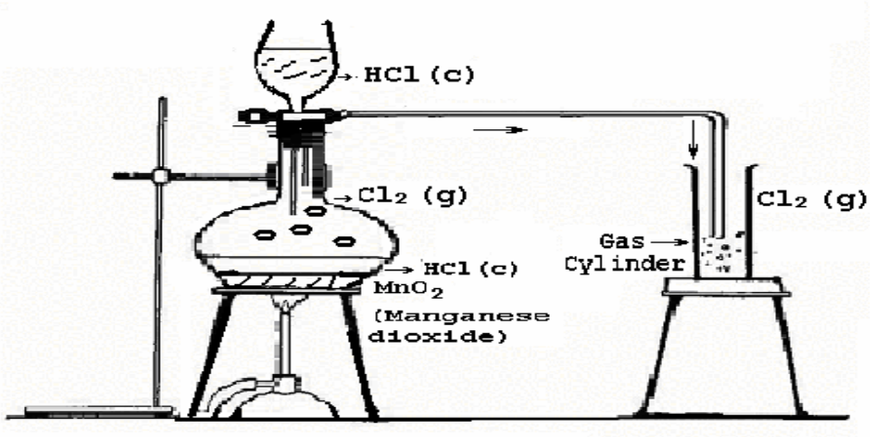
3.3 The preparation of chlorine (Cl_2):

i. In the industry chlorine is prepared by means of electrolyses (sending a current through a salt) as follows:



Nelson Cell

ii. In the laboratory:



Reaction: $4HCl + MnO_2 \xrightarrow{\Delta H} Cl_2 + MnCl_2 + 2H_2O$

Explanation: $Cl^- \longrightarrow Cl_2(g) + 2e^-$

$MnO_2 + 4H^+ + 2e^- \longrightarrow Mn^{2+} + 2H_2O$ (See table)

$Cl^- + MnO_2 + 4H^+ \longrightarrow Cl_2 + Mn^{2+} + 2H_2O$

Now: Balance by ion-inspection.

$4Cl^- + MnO_2 + 4H^+ \longrightarrow Cl_2(g) + Mn^{2+} + 2H_2O + 2Cl^-$

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3.4 Typical reactions of Cl_2:

i. With metals:

$2Na(s) + Cl_2 \longrightarrow 2NaCl$

$2Na(s) + I_2 \longrightarrow 2NaI$

$Mg\ (lint) + I_2 \longrightarrow Mg\ I_2$

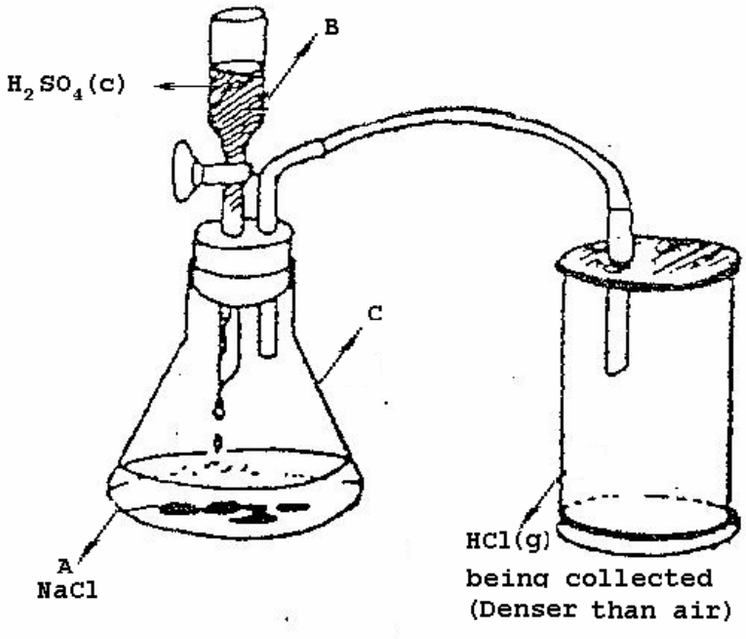
$2Fe(s) + 3Cl_2 \longrightarrow 2Fe\ Cl_3$

ll. Dissolves in water

$Cl_2 + H_2O \longrightarrow HCl + HOCl$

Disinfectant (kills germs in drinking water and swimming pools) + bleaching agent (removes stains).

3.4 The preparation of hydrogen chloride (hydrochloric acid) - HCl:



$H_2SO_4 + 2NaCl \longrightarrow Na_2SO_4 + 2HCl$

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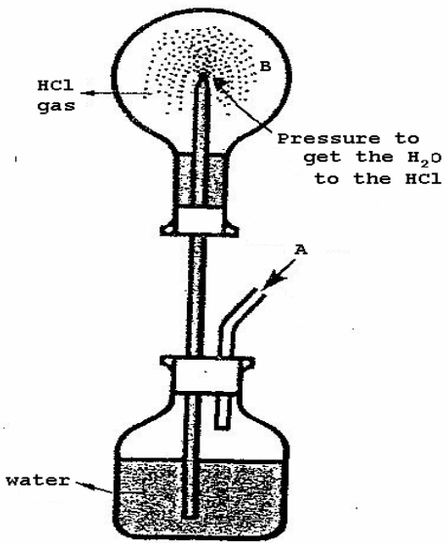
i. Properties of HCl:

- Reaction with H_2O:



$HCl + H_2O(l) \longrightarrow H_3O^+ + Cl^-$

- See acid proton donation of acids
- Litmus turns red.

The fountain experiment illustrates the solubility of HCl in water.

- As reducing agent (in oxidizing reactions)

i. $2Cl^-$ (from HCl) $\longrightarrow Cl_2 + 2e^-$ (oxidation)
(Reducing agent)

$MnO_2 + 4H^+ + 2e^- \longrightarrow Mn^{2+} + 2H_2O$ (See table)

$2Cl^- + MnO_2 + 4H^+ \longrightarrow Cl_2 + Mn^{2+} + 2H_2O$

Balance ions:

$4Cl^- + MnO_2 + 4H^+ \longrightarrow Cl_2 + Mn^{2+} + 2Cl^- + 2H_2O$

- $2Cl^- \longrightarrow Cl_2 + 2e^-$

$Cr^2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 3Cl_2 + 2Cr^{3+} + 7H_2O$

$6Cl^- + Cr_2O_7^{2-} + 14H^+ + \longrightarrow 3Cl_2 + 2Cr^{3+} + 7H_2O$

Now: Balance the $^-$ions.

$14Cl^- + Cr_2O_7^{2-} \longrightarrow 3Cl_2 + 2Cr^{3+} + 7H_2O + 2Cl^-$

- $2Cl^- \longrightarrow Cl_2 + 2e^-$

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$MnO_4^- + 8H^+ 5e^- \longrightarrow Mn^{2+} + 4H_2O$

$10Cl^- + 2MnO_4^- 16H^+ \longrightarrow 5Cl_2 + 2Mn^{2+} + 4H_2O$

Now: Balance **all** the spectator ions in this reaction:

$16Cl^- + 2MnO_4^- + 16H^+ + 2K^+ \longrightarrow 5Cl_2 + 2Mn^{2+} + 2Cl^- + 2K^+ + 4H_2O$

Given!

Given!

ii. Uses of HCl:

- Cleaning in the building industry (of bricks):

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

- Cleaning metal surfaces.
- Prevent algae in swimming pools.

3.5 The halides (salts of the acids of Group 7):

i. Properties of the halides:

Group 1 metal halides (NaCl, KBr, etc) are soluble in water.

Group 2 metals form halides ($MgCl_2$, $CaCl_2$, etc.) which are not easily dissolved in H_2O.

ii. Test for the halides:

A silver nitrate solution is added to the halide solution, forming the following:

$NaCl + AgNO_3 \longrightarrow NaNO_3 + AgCl$ (white)

$KBr + AgNO_3 \longrightarrow KNO_3 + AgBr$ (yellow)

$KI + AgNO_3 \longrightarrow KNO_3 + AgI$ (yellow)

Silver Nitrate

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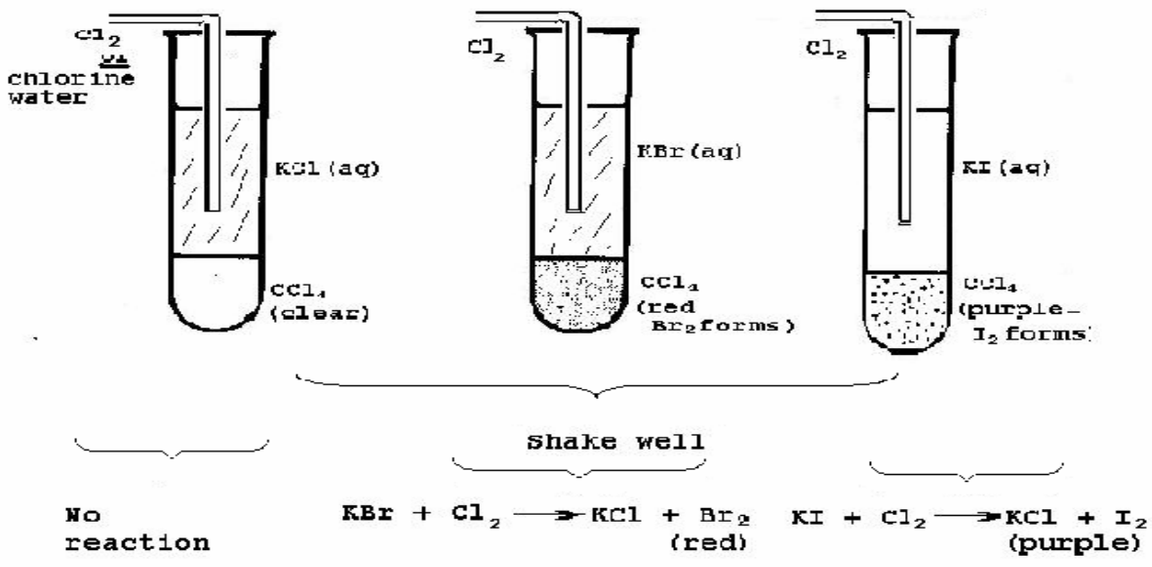
Final Test:



Phases of matter

1. The gas phase (or state:)

1.1 Introduction:

All the different phases of matter can be explained in terms of the "kinetic molecular theory of matter", always pointing to the importance of:

- Intermolecular forces; and
- The motion of particles or molecules.

1.2 The kinetic theory:

The following concepts is important:

- Real gasses (non-ideal)
- Ideal gasses (non-real)

"An ideal gas is a gas which will predict and explain all the characteristics and properties of gasses."

"A real gas deviates from this ideal behaviour of the 'ideal gas' because of mainly the following:

- Gas molecules occupies volume due to its molecules (especially at very low temperatures and very high pressures).

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- Gas molecules have forces between them (especially at very low temperature and very high pressures).

- The collisions between the molecules are not fully elastic.

The ideal gas obeys "Boyle's Law" because of the following assumptions for an ideal gas (real gasses "also obeys 'Boyle's Law' at STP-standard temperature and pressure, eg. 0°C and 101,3 kPa):

- The gas molecules are far apart - they exert no force of attraction on one another except when they collide (they fill the container in which they occur).

- The molecules of a gas are in a state of constant motion and they collide with each other and with the walls of the container. The collisions of the molecules of the gas with the walls of the container, is called the pressure of the gas.

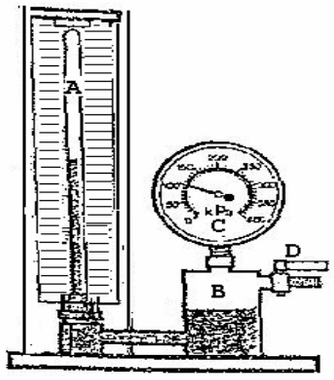
- The temperature of a gas is the measurement of the average kinetic energy of its molecules. (Heat of a gas or object is the amount of energy absorbed or given of).

- Diffusion (where the molecules of one gas moves into the open spaces of another gas) to form a mixture of gasses, easily takes place.

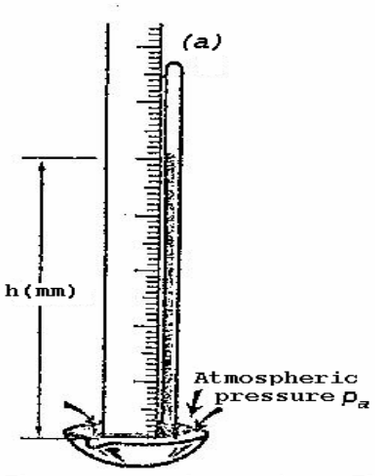
- All the collisions of particles are perfectly elastic, therefore no energy is lost during the collision.

1.3 Boyle's Law:

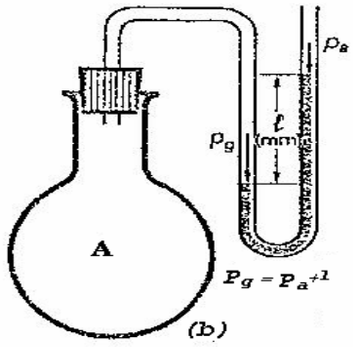
1.3.1 Measuring pressure and volume:



The apparatus used to investigate the relationship between the pressure and the volume of an eclosed gas as constant temperature.



The mercury barometer. The length of the mercury column (h) is converted into atmospheric pressure p_a in kPa (7,5 mm Hg=1 kPa).



A manometr is used to measure the pressure exerted by a gas

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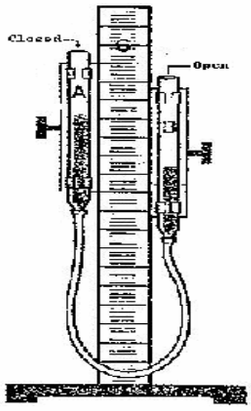
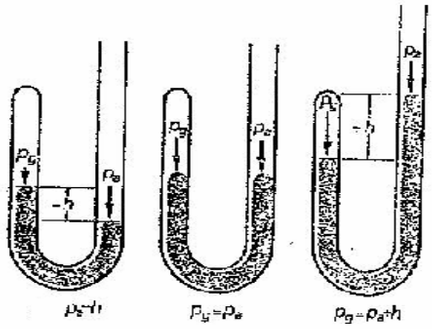
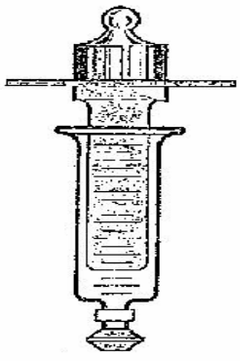


The apparatus used to investigate the relationship between the pressure and the volume of a quantity of gas



P_g = Pressure of gas
P_a = Pressure of atmosphere
h = difference in levels of mercury columns (mm) expressed in pressure unit (kPa).



A gas syringe may be used to investigate the pressure-volume relationship

In all the cases shown above, a certain volume of gas is brought under a certain pressure to determine the relationship (Boyle's Law):

Typical results can be as follows:

Volume (V) of gas in cm^3	Pressure (p) on gas in kPa	**P1V1 / P2** in cm^{-3}	pV (constant)
15	200	$6,7 \times 10^{-2}$	$3x10^3$
20	150	$5,0 \times 10^{-2}$	$3x10^3$
23	130	$4,3 \times 10^{-2}$	$3x10^3$
30	100	$3,3 \times 10^{-2}$	$3x10^3$
35	85,5	$2,9 \times 10^{-2}$	$3x10^3$
42	71,5	$2,4 \times 10^{-2}$	$3x10^3$
50	60	$2,0 \times 10^{-2}$	$3x10^3$

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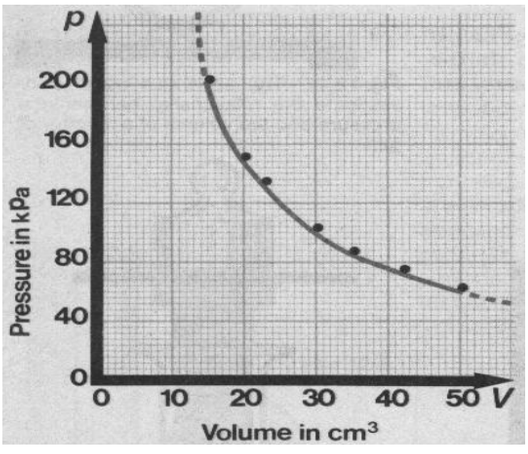

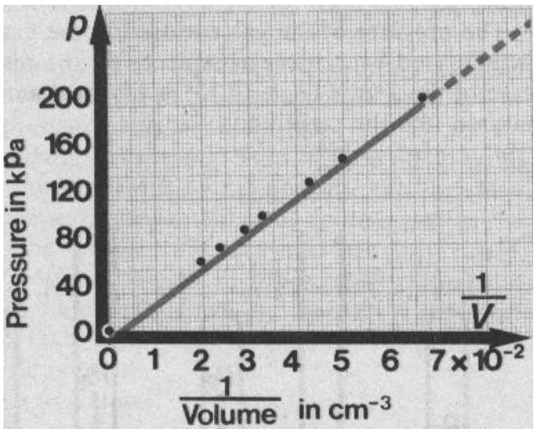

SG Examples:

i. The volume of air in a balloon is 2000 cm^3 at a pressure of 80 kPa (temp. constant). What will the volume of the balloon be at 60 kPa?

Data: $V_1 = 2000\ cm^3$; $V_2 = ?$

$P_1 = 80\ kPa$; $P_2 = 60\ kPa$

Now: $P_1V_1 = P_2V_2$

$P_2V_2 = P_1V_1$

$$V_2 = \frac{P_1V_1}{P_2}$$

$$= \frac{80 \times 2000}{60}$$

$$= 2667 cm^3$$

The pressure decreases, therefore the volume increases (indirectly proportional to each other).

SG At what pressure will the balloon burst if its maximum volume can only be 5000 cm^3?

Data: $P_1 = 80\ kPa$; $P_2 = ?$

$V_1 = 2000\ cm^3$; $V_2 = 5000\ cm^3$

Now: $P_1V_1 = P_2V_2$

$P_2V_2 = P_1V_1$

$$P_2 = \frac{P_1V_1}{V_2}$$

$$= \frac{80 \times 2000}{5000}$$

$$= 32 kPa$$

HG ii. A car tyre has a volume of 2,5 x 10^{-2} m^3 of air at atmospheric pressure of (say) 100 kPa, is pumped into the tyre determine the pressure in the tyre (temp. constant).

Data: $P_1 = 100\ kPa$; $P_2 = ?$

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$V_1 = 5,0 \times 10^{-3}\ m^2$ $V_2 = 2,5 \times 10^{-2}\ m^3$

Now: $P_1V_1 = P_2V_2$

$\Rightarrow$ $P_2V_2 = P_1V_1$

$\Rightarrow$ $P_2 = \frac{P_1V_1}{V_2}$

$$\frac{100 \times 5,0 \times 10^{-2}}{2,5 \times 10^{-2}}$$

$= 200$ kPa

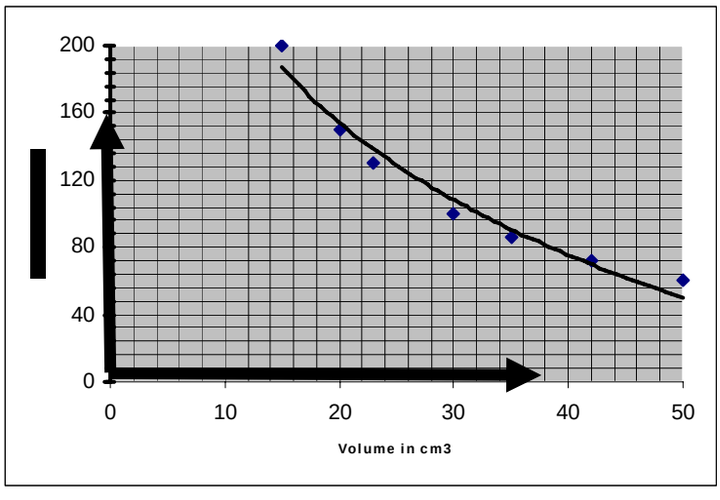
SG iii.

p (kPa)	V(cm^3)
180	25
300	15
450	10

a) State (in words) the mathematical relationship between P and V.

The volume of a gas is inversely proportional to the pressure on it (if the temperature remains constant).

b. Graph: c. State one factor that has to be constant temperature



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d. V_2 if $P_2 = 200$ kPa

Now: $P_1V_1 = P_2V_2$

$\Rightarrow \quad P_2V_2 = P_1V_1$

$$\Rightarrow \quad V_2 = \frac{P_1V_1}{P_2}$$

$$= \frac{180 \times 25}{200}$$

$$= \frac{4500}{200}$$

$$= 22,5\text{cm}^3$$

e. Very low temperatures and very high pressures, the relationship will no longer be valid (see gas molecular theory).

1.4 Explanation of Boyle's Law: (in terms of the molecular kinetic theory for gasses): The volume of an enclosed gas is inversely proportional to the pressure exerted on it, if the temperature remains constant. $P_1V_1 = P_2V_2$ (Temp. constant)

Explained as follows:

- The temperature remains constant ⇒ the average kinetic energy of the molecules remains constant.
- When the volume of the gas is decreased, the available "space" for the movement of the molecules decreases, more collisions will take place per time and the pressure increases. (Remember the collisions on the wall of the container are defined as the pressure of the gas).
- When the volume of the confined gas increases, the available "space" for the movement of the gas molecules increases, less collisions will take place per time and the pressure decreases (temp. remains constant).

Read only:

1.5 The Kelvin (K) temperature scale:

From the fact that pV&T, the Kelvin scale was derived taking into account that the upper point of the Kelvin scale is 273,16 K (or 0°), where ice, water and water vapour are in equilibrium at the same time (the so-called triple point of water). The absolute zero (0°K **or** -273.16°C) cannot be reached. Note: 1K = 1°C, therefore 20°C is equal to:

$$T = 273 + °C = 273+20$$

$$= 293 \text{ K (above or higher than)}$$

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1.6 Other derived formulas (and Boyle's Law:

$P_1V_1=P_2V_2$ (Temperature constant)	$\frac{V_1}{T_1}=\frac{V_2}{T_2}$ (Pressure constant)	$\frac{P_1}{P_2}=\frac{T_1}{T_2}$ (Volume constant)	$\frac{P_1V_1}{T_1}=\frac{P_2V_2}{T_2}$ (Everything changes)	$pV=nRT$ where n=number of moles R=General gas constant

and

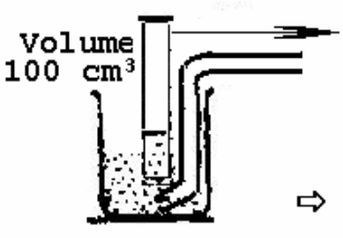
$P_{total} = P_1 + P_2 + P_2$

The total pressure of a mixture of gasses is the sum of the separate (partial) pressure of the individual gasses.

1.7 Examples:

1.7.1

a.



Ptotal = 140kPa
PH_2O (g) = 3,0 kPa
O_2 Po_2 = ?
Temp. 25°C

Ptotal = P_1 P_2
⇨ Po_2 = Ptotal - PH_2O (g)
= 140 - 3
= 137kPa

Find the volume of dry oxygen at STP?

P_1 = 137 kPa V_1 = 100 cm^3 T_1 = (273 + 25) = 298 K →	P_2 = 101,3 kPa V_2 = ? T_2 = (273 + 0) = 298 K →

$$\frac{P_1V_1}{T_1}=\frac{P_2V_2}{T_2}$$

$$\Rightarrow P_2V_2T_1 = P_1V_1T_2$$

$$\Rightarrow V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

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$$= \frac{137 \times 100 \times 123}{101,3 \times 298}$$

$$= \frac{3740100}{30187,4}$$

$$= 123,9\ cm^3$$

SG & HG

1.7.2 A gas is collected in a 500 ml flask at 25°C at a existing pressure of 150 kPa. Find the number of moles of gas in the flask.

Now: $pV = nRT$

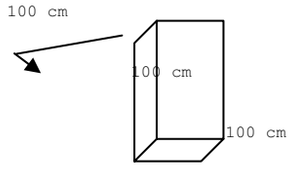
$P = 150\ kPa = 150\ 000\ Pa$

$V = 500\ ml = 500 \times 10^{-6}\ m^3$

$n = ?$

$R = 8,31\ J/K.mol$

$T = 273 + 25 = 298K$



$$pV = nRT$$

$$\Rightarrow \quad nRT = pV$$

$$\Rightarrow \quad n = \frac{pV}{RT}$$

$$= \frac{150\ 000 \times 500\ 10\ (-6)}{8.31 \times 298}$$

$$= 0.03\ mol$$

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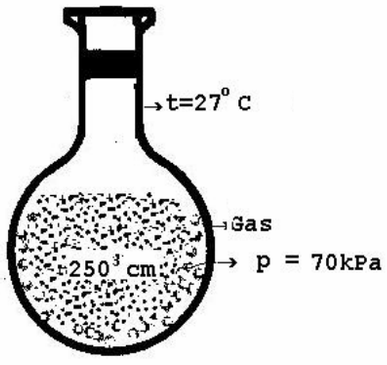
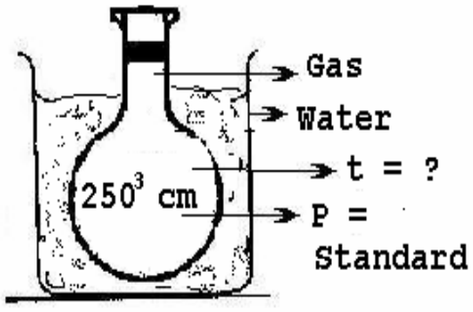
1.7.3





Data:

Before:

$P_1 = 70$ kPa

$T_1 = 273 + 27$

$= 300$ k

$V_1 =$ konstant

After:

$P_2 =$ Standard

$= 101,3$ kPa

$T_2 = ?$

$t_2 = ?$

$V_2 =$ Konstant

$$\frac{P_1}{P_2} = \frac{T_1}{T_2}$$

$$\Rightarrow \quad T_2 p_1 = T_1 p_2$$

$$\Rightarrow T_2 = \frac{T_1 p_2}{P_1}$$

$$= \frac{300 \times 101,3}{70}$$

$$= 434 \text{ K}$$

But: $T_2 = 273 + t_2$

$$\Rightarrow \quad t_2 = T_2 - 273$$

$$= 434 - 273$$

$$= 161°C$$

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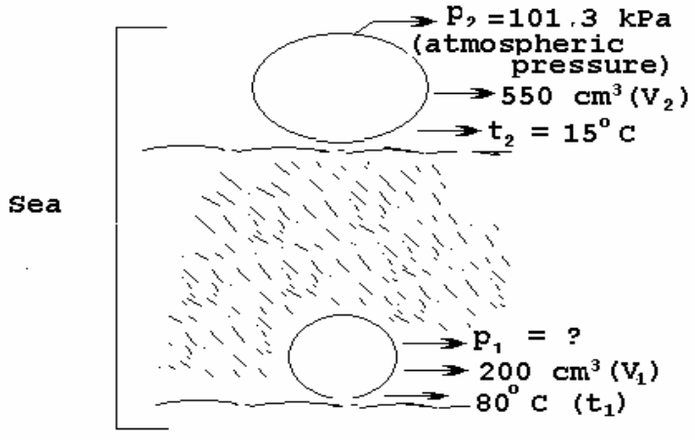
1.7.4 A gas bubble deep below the surface of the sea has a volume of 200 cm^3 at 8°C. The bubble rises to the surface where t_2 = 15°C and the volume increases to 550 cm^3. Find p at the original.



$$\frac{P_1V_1}{T_1} = \frac{p_1V_2}{T_2}$$

$$\therefore P_1V_1T_2 = P_2V_2T_1$$

$$\therefore P_1 = \frac{p_2V_2T_1}{V_1\ T_2}$$

$$= \frac{101,3 \times 550 \times (273 + 8)}{200 \times (273 + 15)}$$

$$= \frac{101,3 \times 550 \times 281}{200 \times 288}$$

$$= \frac{15655915}{57600}$$

$$= 271,8kPa \rightarrow$$

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2.3 The Bronsted-Lowry Theory of acids and bases:

2.3.1 Acids are defined as proton (H^+) donors; and

2.3.2 bases are defined as proton (H^+)acceptors.

2.3.3 In neutralization of an acid with a base, called protontransfer or protolyses a proton ((H^+)is transferred from the acid (proton donor) to the base (proton acceptor or receiver):

i. $HCl + NH_3 \rightleftharpoons NH_4^+ + Cl^-$

Acid (Donor) 1 | Base (Receiver) 2 | Acid 2 | Base 1

Acid-base Pair Conjugate Pair 1

Conjugate acid-base Pair 2

ii. $HCN + H_2O \rightleftharpoons H_3O^+ + CN^-$

Acid 1 | Base 2 | Acid 2 | Base 1

2.3.4 More protolytic reactions:

$H_2SO_4 + H_2O \rightleftharpoons H_3O^+ + HSO_4^-$

$NH_4^+ + F^- \rightleftharpoons HF + NH_3$

$H_3O^+ + OH^- \rightleftharpoons H_2O + H_2O$

$H_2O + CH_3COO^- \rightleftharpoons CH_3COOH + OH^-$

$NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH_3$

$H_2SO_4 + Cl^- \rightleftharpoons HCl + HSO_4^-$

2.3.5 Amphiprotic substance (protobytes:

Substances such as HSO_4^- can act as an acid or a base, are called protobytes:

$HSO_4^- + H_3O^+ \rightleftharpoons H_2SO_4 + H_2O$

(Base)

$H_2SO_4 + OH^- \rightleftharpoons SO_4^{2-} + H_2O$

(Acid)

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2.3.6 Hydrolysis (add water to salts) to break the salts up:

i. $H_2O + CO_3^{2-} \rightleftharpoons HCO_3^- + OH^-$

OH^- - ions in solution makes the litmus turns blue

⇒ acid.

Conclusion: Hydrolyses forms acidic or basic solutions ⇒ depending on the type of salt used.

3. Ionization of acids/bases (relative strengths) of acids/bases:

Chemical equilibrium supplies us with a method to compare the relative strengths of acids/bases as follows:

Acids: i) $HCl(g) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$

The $[H_3O^+]$ is an indication of the strength of the acid.

$$\therefore \quad Ks = \frac{[H_3O^+][Cl^-]}{[HCl]}$$

Ks = dissosiation Constant.
$[H_2O]$ in liquid phase.

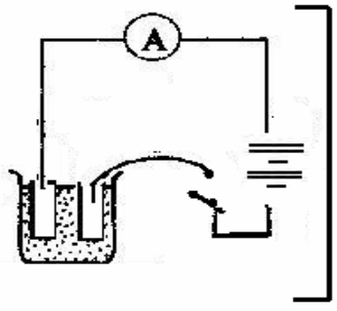
ii) HCl is a relative strong acid because Ks is large, showing that the products are of high concentration.

iii) $CH_3COO^+ + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

$$\therefore \quad Ks = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

iv) Ks is very small, indicating that low concentrations of products are formed ⇒ acetic acid is therefore a very weak acid (1% dissociation)

v) Electrical conductivity as an indicator of the relative strength of acids:



Strong acids conducts electricity <u>well</u>.
Weak acids conducts electricity <u>poorly</u>.

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Bases: i. Strong bases:

$NaOH(s) \rightleftharpoons Na^+(aq) + OH^-(aq)$

$Kb = [Na^+][OH^-]$ = Very large

ii. Weak bases:

$NH_3(q) + H_2O(l) \rightleftharpoons NH_4(q) + OH^-(aq)$

$$Kb = \frac{[NH_4^+][OH^-]}{[NH_3]} = \text{Very small}$$

4. The ionization of water (auto-protolyses):

$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(q) + OH^-(aq) + OH^-(aq)$

$Kur = [H_3O^+][OH^-] = 1,0 \times 10^{-14}$ at 25°⇒ the ionic product is very small.

$[H_3O^+][OH^-] = 1,0 \times 10^{-14}$

$[H_3O^+] = 1,0 \times 10^{-7}$ $\qquad$ $[OH^-] = 1,0 \times 10^{-7}$

Neutral solution

In an acidic solution $\qquad$ In a basic solution

$[H_3O^+] > 1,0 \times 10^{-7}$ $\qquad$ $[HO^-] < 1,0 \times 10^{-7}$

5. The pH - scale:

The pH - scale indicates the degree of acidity:

$$pH = \text{Log} \frac{1}{[H_3O^+]} = -\log [H_3O^+]$$

Examples:

i. In a neutral solution, $[H_3O^+] = 1,0 \times 10^{-7}$

$pH = -\text{Log}\ [H_3O^+]$

$= -\text{Log}\ (1,0 \times 10^{-7})$

$= -\ (0^{-7})$

$= +\ 7$ (neutral) ⟶

ii. In an acidic solution, $[H_3O^+] = 1,0 \times 10^{-4}$

$pH = -\text{Log}\ [H_3O^+]$

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$= - \text{Log } (1,0 \times 10^{-4})$

$= - (0^{-4})$

$= + 4$ (acidic)

iii. In a basic solution, $[H_3O^+] = 1,0 \times 10^{-10}$

$pH = - \text{Log } [H_3O^+]$

$= - \text{Log } (1,0 \times 10^{-10})$

$= - (0^{-10})$

$= + 10$ (Basic)

iv. The $[H_3O^+] = 3,0 \times 10^{-3}$ of a solution. What is the pH?

$pH = - \text{Log } [H_3O^+]$

$= - (\text{Log } (3,0 \times 10^{-3})$

$= - (0,48 + (-3))$

$= 2,52$ (Acidic)

pH The scale of pH-values is as follows:

	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
$[H_3O^+]$	10^{-0}	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
$[OH^-]$	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	10^{-0}

Acidic (Litmus (a plant extract) turns red).

Neutral

Basic (Litmus turns blue).

Examples (HG):

1. Given pH of solution = 6,4. Find $[H_3O^+]$ and $[OH^-]$

$pH = 6,4$

$\therefore \quad pH = - \text{Log } [H_3O^+]$ (Definition)

$\therefore - \text{Log } [H_3O^+] = 6,4$

$\therefore \quad \text{Log } [H_3O^+] = -6,4$

$= + 0,6 - 7,0$

$\therefore \quad [H_3O^+] = \text{Antilog } 0,6 \times \text{Antilog } (-7)$

$= 3,98 \times 10^{-7}$ mol/dm^3 (mol/l)

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$\therefore [H_3O^+][OH^-] = 1,0 \times 10^{-14}$

$\therefore [OH^-] = \frac{1,0 \times 10^{-14}}{3,98 \times 10^{-7}}$

$= 2,51 \times 10^{-8}$ mol/dm^3 →

2. If Kc = $1,8 \times 10^{-5c}$ at 25°C for ethanoic acid, find the pH if $[CH_3COOH] = 0,15$ mol/dm^3

$CH_3COOH + H_2O(l) \rightleftharpoons H_3O^+ + CH_3COO^-$

(99% of acid don't ionise)

$\therefore Kc = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$

Let $[H_3O^+] = [CH_3COO^-] = X$

$\therefore Kc \times [CH_3COOH] = X^2$

$X^2 = 1,8 \times 10^{-5} \times 0,15$

$= 2,7 \times 10^{-6}$

$\therefore X = 1,64 \times 10^{-3}$

$\therefore$ But $x = [H_3O^+]$

$\therefore [H_3O^+] = 1,64 \times 10^{-3}$ mol/dm^3 →

$pH = - \text{Log } [H_3O^+]$

$= - \text{Log } (1,64 \times 10^{-3})$

$= - 0,22 + 3$

$= 2,8$ (acidic) →

The measurement of pH in the laboratory:

pH-meter	Or	Universal Indicator

Examples (using the formula):

i. In a titration 65 ml of NaOH-solution of concentration 0,12 mol/dm^3 neutralises 50 ml of HNO_3.

HG

$\underbrace{NaOH}_{1\text{mole}} + \underbrace{HNO_3}_{1\text{mole}} \rightarrow \underbrace{NaNO_3}_{1\text{mole}} + \underbrace{H_2O}_{1\text{mole}}$

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$$\frac{Ca\ Va}{Cb\ Vb} = \frac{Na}{Nb}$$ **(See formula sheet)**

$\therefore$ **Na Cb Vb = Ca Va Nb**

= 0,0156 mol/dm^3

SG ii.

30 cm^3 of H_2SO_4(dil) neutralises 20cm^3 of a 2,0 mol/dm^3 solution of Mg $(OH)_2$.

Find $[H_2SO_4]$?

Reaction: $Mg(OH)_2 + H_2SO_4(aq) \rightarrow 2H_2(l) + Mg(SO_4(aq)$

(Given)

$\frac{Ca\ Va}{Cb\ Vb} \not= \frac{Na}{nb}$ **(See information sheet)**

$\therefore$ **Ca Va nb = cb Vb na**

$\therefore$ **Ca** $= \frac{cb\ Vb\ na}{Va\ nb}$

$= \frac{2,0 \times 20 \times 1}{30 \times 1}$

$= \frac{40}{30}$

= 1,33 mol/dm^3

Redox reaction and Electrochemistry

1. Introduction:

1.1 Oxidation takes place when an atom gives of electrons; and

1.2 Reduction takes place when an atom takes up electrons.

1.3 Oxidation and reduction takes place at the same time in a reaction - therefore the reactions are called redox-reactions.

1.4 In the oxidation reaction, we find the reducing agent (the reactant causing the oxidation), eg.

$H_2S \rightarrow S + 2H^+ + 2e^-$ (oxidation)

Reducing agent

1.5 $MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$

(Reduction)

2. Electron transfer (takes place at the same time - redox reaction)

2.1 **Mg + O_2 $\rightarrow$ MgO, can be broken down into:**

$Mg \rightarrow Mg^{2+} + 2e^-$ (Oxidation)

$O + 2e^- \rightarrow O^{2-}$ (*Re*duction)

$\therefore$ **Mg + O $\rightarrow$ MgO**

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2.2 **$Mg + Cl_2 \rightarrow MgCl_2$, can be broken down into:**

$Mg \rightarrow Mg^{2+} + 2e^- \ (Oxidation)$

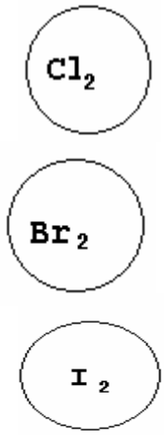
$Cl_2 + 2e^- \rightarrow 2Cl^- \ (Reduction)$ reducing agent

Remark: Although the reaction is split up into two half-reactions, they are added up to show the whole redox-reaction.

3. Reactivity in terms of "the ease of donating" electrons:

3.1 Introduction:

Group 7: (The halogens)



Chlorine Is "relatively Small" and takes Up e^- easily, eg. From Br_h.

$2KBr + Cl_2(g) \rightarrow 2KCl + Br_2$

$2KI + Cl_2(g) \rightarrow 2KCl + I_2$

Br is "smaller" than I_2 and takes up Electrons from I_2

$KCl + Br_2 \rightarrow$ No reaction

$2KI + Br_2 \rightarrow 2KBr + I_2$

The strength of the oxidizing agents relates as follows:

$Cl_2 > Br_2 > I_2$

eg. $Cl_2 + 2e^- \rightarrow 2Cl^-$ (Reduction)

(Oxidizing agent)

The "donating electrons", oxidation, increases from:

$I_2 > Br_2 > Cl_2$

$I_2 \rightarrow 2I^- + 2e^-$ (Oxidation)

(Reducing agent)

The strength of the reducing agents relates as follows:

$I_2 > Br_2 > Cl_2$

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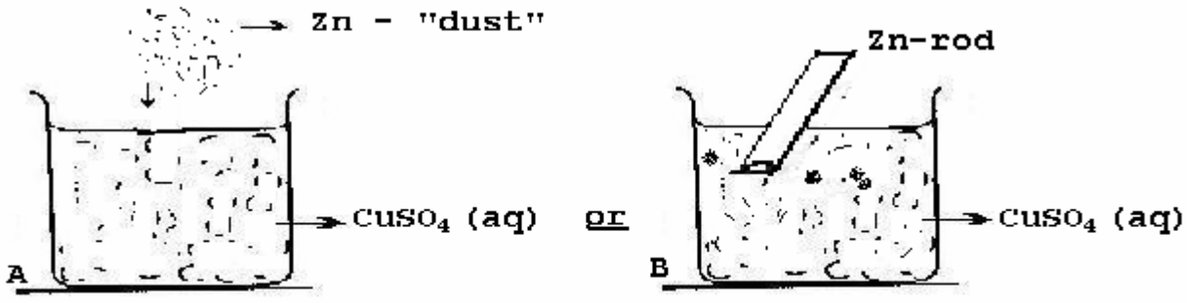
4. Redox reactions: (Direct transfer of e^-)

4.1



The solution of $CuSO_4$ discolours (left) and on the Zn-rod (left) a red deposit forms. After a while a red presipitate can be seen at the bottom of the beaker A.

Tests:

i. Red deposit/presipitate:

$$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$$

(Blue)

ii. Proof: Cu is deposited.

i. $$H_2S(g) + ZnSO4 \rightarrow ZnS\downarrow + H_2SO_4$$

White Precipitate

ii. Proof: **Zn^{2+} ions in solution. (This is the test for Zn^{2+} – ions).**

From this tests: Cu is deposited (from the solution) and Zn disappears from the "dust" or rod **(into the solution as Zn^{2+} - ions).**

$Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation)

$Cu^{2+} + 2e^- \rightarrow Cu$ (reduction)

$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$ (Redox)

4.2 Indirect transfer of electrons (redox reactions in electro-chemical cells):

4.2.1 The facts:

a. The two reactions are physically separated (oxidation and reduction occurs separately).

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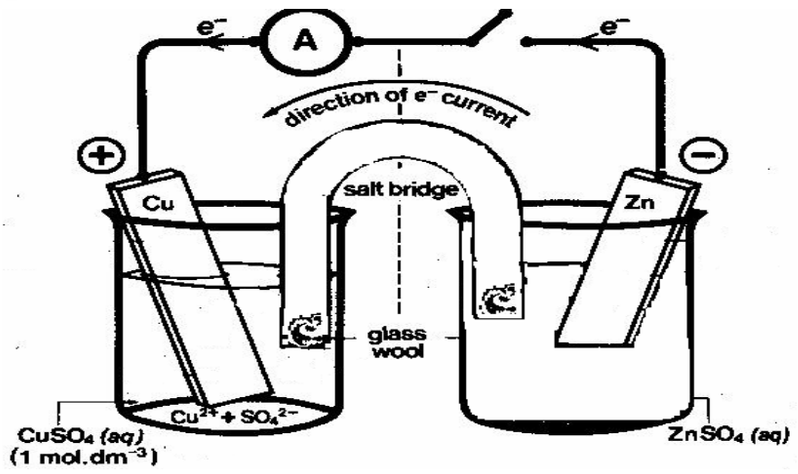
b. The flow of electrons can be detected, used and or measured.

c. The transfer of electrons takes place indirectly through a conductor externally.

4.2.2 The basic cell: The Cu/Zn electro-chemical cell.

i.



ii. Cell notation:

- $Zn(s)/Zn^{2+}\left(1mol/dm^3\right)//Cu^{2+}\left(1mol/dm^3\right)/Cu(s)^{\oplus}$ +

iii. Working of the cell:

$$Zn(s) \rightarrow Zn^{2+} + 2e^{-}\,(oxidation)$$

(Electrons pile up on the Zn and the pole of the cell becomes the negative terminal - ANODE)

$$Cu^{2+} + 2e^{-}\ Cu(s)\downarrow(reduction)$$

(Electrons are taken up on the Cu-electrode [pole] of the cell by

Cu^{2+} - ions and this deficiency is causing the pole to become positive - CATHODE).

iv. Total cell reaction. Add up):

$$Zn(s) + Cu^{2+} \rightarrow \underbrace{Zn^{2+}} + \underbrace{Cu(s)}\downarrow$$

The Zn-electrode dissolves

The Cu^{2+} takes up e^- and Cu is deposited on the Cu-electrode

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5. Cell potential

HG 5.1 The cell potential is the ability of the cell to do work.

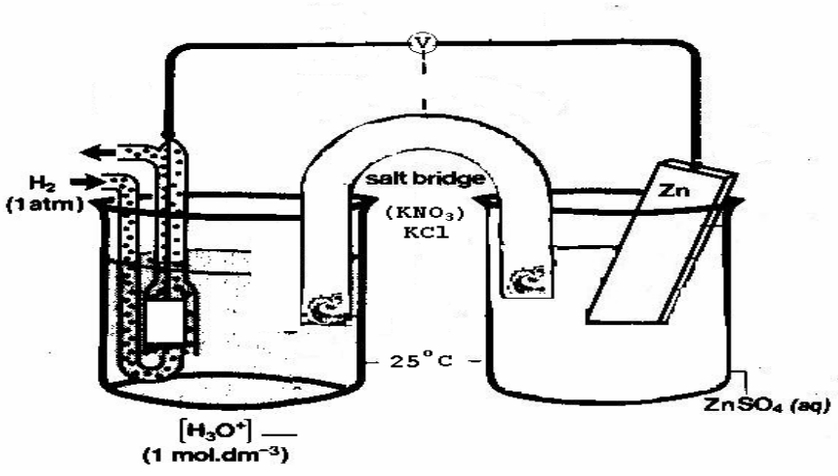
5.2 The standard cell conditions:

Standard conditions:

Temp: 25°C Leads to

Concentration: /mol/dm^3 } Standard Emf.

5.3 The standard Cell:



The left cell is a standard hydrogen electro with a potential of 0,00 Volt.

All the metals can be compared with this standard hydrogen electrode and the table of electro potentials can be compiled.

$$(-)\ \square\ Pt, H_2 / H^+ \left(1mol/dm^3\right) // Cu^{2+} \left(1mol/dm^3\right) / Cu^{\oplus}\ (+)$$

Results for (say Cu):

$$Cu^{2+} + 2e^- \rightarrow Cu(s);\ E^{\square} = +0,34V$$

All the different half-reactions can, in the same way, be comp[ared with the standard hydrogen half-cell and a table set up.

5.4 Examples:

SG 5.4.1 $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+} + Cu(s)$

Break up in half-cells:

$$Zn \rightarrow Zn^{2+} + 2e^-$$

$$Cu^{2+} + 2e^- \rightarrow Cu(s)$$

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Can Zn donate electrons to the Cu^{2+} - ions?

See table of electrode standard potentials:

$$Zn \rightarrow Zn^{2+} + 2e^- \qquad \left(E^{\Theta} = -0,76V\right)(ANODE)$$

$$Cu^{2+} + 2e^- \rightarrow Cu \qquad \left(E^{\Theta} = +0,34V\right)(CATHODE)$$

$$E_{cell}^{\Theta} = E^{\Theta}_{\text{oxidising agent}} - E^{\Theta}_{\text{reducing agent}}$$

HG $$E_{cell}^{\Theta} = E_{Cathode}^{\Theta} - E_{Anode}^{\Theta}$$

$$= +0,34 - (-0,76)$$

$$= 1,10V \ (\text{voltmeter reading})$$

The answer is positive ⇒ the reaction can take as given (a spontaneous reaction).

5.2.4

Will the following proceed spontaneously from the left to the right or from the right to the left?

$$2Ag(s) + Cu^{2+}(aq) \rightarrow Cu(s) + 2Ag^+$$

Implies the following

$$2Ag \rightarrow 2Ag^+ + 2e^-$$

$$Cu^{2+} + 2e^- \rightarrow Cu$$

See table: Because of the position of Ag and Cu wrt oxidising ability, Ag cannot donate e^- to Cu^{2+}. Therefore the reaction as shown cannot take place. Calculation also show that $E^{\Theta}_{cell} \rightarrow = -0,46$ reaction cannot take place as shown (second prove).

$$Cu \rightarrow Cu^{2+} + 2e^- \text{ (oxidation).}$$

$$2Ag^+ + 2e^- \rightarrow Ag \quad (\text{reduction})$$

$E_{cell}^{\Theta} = 0,46V$ (the reaction will be as follows:

$$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag\downarrow)$$

SG Example 5.4.3:

Will the following reaction takes place spontaneously?

$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)\uparrow$$

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Cell half-reactions:

$$Zn \rightarrow Zn^{2+}(aq) + 2e \quad \textbf{(oxidation)}$$

$$2H^{+} + 2e^{-} \rightarrow H_2(g) \quad (\textbf{Reduction})$$

The table and $E^{\ominus}_{cell} = +0{,}76V$ shows that the reaction can take place as shown.

Example 5.4.4:

Which one of the following reactions is a redox-reaction:

i. $NaCl(s) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$

Ion Exchange

ii. $NO_3^{-}(aq) + H_2SO_4(aq) \rightarrow H_2SO_4^{-}(aq)\ HNO_3(aq)$

Protolyses = acid/base reaction

iii. $NH_3(g) + H_2O(l) \rightarrow NH_4^{+}(aq) + OH^{-}(aq)$

Hydrolyses (adding of water)

iv. $Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + NO_2 + 2H_2O$

Giving off of electrons

Redox (see table for HNO_3 reaction).

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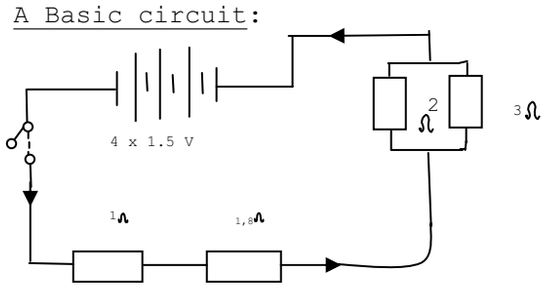
CURRENT ELECTRICITY: ELECTRICAL CIRCUITS

1. A Basic circuit:



1.1 Find the effective resistance of the parallel resistors.

$$\frac{1}{R_p} = \frac{1}{R1} + \frac{1}{R2}$$

$$= \frac{1}{2} + \frac{1}{3}$$

$$= \frac{3}{6} + \frac{2}{6} \quad \text{(Find the smallest common denominator)}$$

$$= \frac{5}{6} \quad \text{(The denominators are the same – add up)}$$

$$R_p = \frac{6}{5} \quad \text{(One single term on both side allows us to invert both sides)}$$

$$= 1{,}2\ \Omega \quad \text{(This single value represents the parallel resistors)}$$

1.2 Find the total effective resistance of the external circuit.

$$R_s = R1 + R2 + R3 = 1{,}0 + 1{,}8 + 1{,}2 = 4.0\Omega$$

Note: Remember the ammeters are regarded as having no resistance in this basic course.

1.3 Find the main current in the circuit.

$$I = \frac{V}{R}$$

Remember this is Ohm's Law) – The current in a circuit is directly proportional to the potential difference across a component and inversely proportional to the resistance of the component.

$$= \frac{(4 \times 1{,}5)}{4}$$

There are 4 cells of 1,5Vin series in the battery.

$$= \frac{6}{4} = 1{,}5\ \text{A (ampere)}$$

Note: The current is 1,5 A everywhere in the circuit – except where branching takes place:

At the parallel connection.

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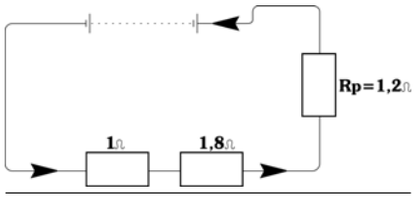
1.4 Find the currents in the branches of the parallel connection.
First draw a equivalent circuit as follows:



1.4.1 Now, we calculate the pd (potential difference) across the parallel connected resistors:

$$V = I \times R = 1,5 \times 1,2 = 1,8 \text{ volt (V)}$$

(Note: Remember the main current flows through the effective resistance (Rp) of the parallel connected resistors with value 1.2 ohm).

(Check the pd's across all the resistors:
V1= 1,5 volt (V); V2= 2.7 volt (V).
Vt = 1,5 + 2,7 + 1,8 = 6 volt (V) = 6 V of the battery).
Remember: Serie resistors are "voltage dividors")

1.4.2 Now, the currents in the branches can be determined easily:

$$I_1 = \frac{V}{R_1} = \frac{1,8}{2} = 0,9 \text{ A}$$

(This is the current branch 1)

$$I_2 = \frac{V}{R_2} = \frac{1,8}{3} = 0,6 \text{ A (current branch 2)}$$

Remember: A parallel connection of resistors is a "current divider", and therefore

$$I_{main} = I_1 + I_2 = 0,9 + 0,6 = 1,5A$$

(See previous calculation).

1.5 Determine the amount of heat that will develop in the 1,8 ohm (Ω) if the main current (remember that the main current flows everywhere in the circuit except when the circuit "splits up", for example in the Rp) flows for 1 minute and 32 seconds.

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(Formulas: $W = I^2Rt$ or $W = VIt$ or $W = \frac{V^2t}{R}$)

Note: W represents heat, work done or energy and it is measured in joule (J).

$W = I^2Rt = (1,5 \times 1,5) \times 1.8 \times (60 + 32) = 2,25 \times 1,8 \times 92 = 372,6$ joule (J) of energy (heat) is dissipated.

(Note:

- Remember that any other of the formulas could have been used – it requires an extra calculation).
- Any other component could have been selected – remember to use only information
- applicable to that specific component; forget about the other information available if it is not required here.

1.6 Determine the power consumed in the 3 ohm resistor if the current flows through it for 1 minute and 32 seconds.

(Formulas: $P = VI$ or $P = I^2R$ or $P = \frac{V^2}{R}$)

$P = VI = 1,8 \times 0,6 = 1,08$ watt (W) meaning that 1,08 joule (J) of electrical energy is transformed into heat per second.

(Note:

- Time given is not used in this calculation of power.
- The time can be used to determine the total amount of energy being dissipated in the 92 s (s = seconds), namely in the 92 s, 92 x 1,08 or 99,36 joule (J) of heat developes).

1.7 Determine the amount of charge that moves through the 1 ohm resistance if the main current is allowed to flow for 1 minute and 53 seconds.

$Q = It$ (This formula is given and it also allows us to define current, Q

$= 1,5 \times 113$ namely $I = \frac{Q}{t}$,

Giving us 1A of current flows when 1 C of charge passes from the one point to the other

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1.8 How many electrons move pass a point in the resistor of 1 ohm in 1 minute and 53s?

$$\text{Number of electrons} = \frac{\text{Total charge passing the point}}{\text{Charge on a single electron}}$$

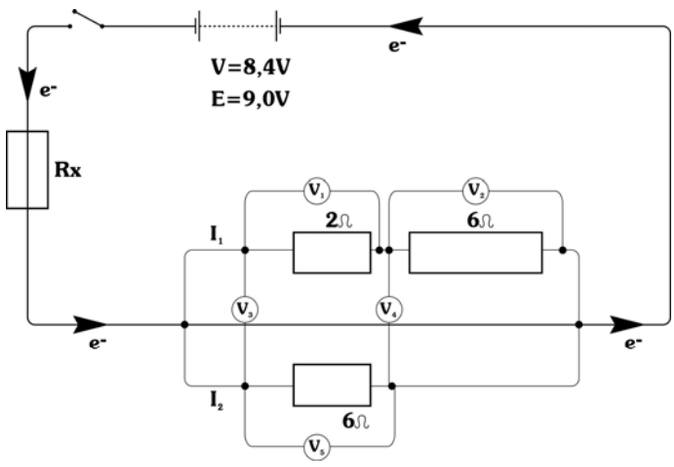
$$= \frac{169,5}{1,6 \times 10^{-19}}$$

$$= 105,938 \times 10^{19}$$

$$= 1,06 \times (10^{2} \times 10^{19})$$

$$= \underline{1,06 \times 10^{21}}$$

2. The following circuit is applicable to the higher grade:



For the HG the battery is regarded as a component also having resistance, called the internal resistance of the battery (r). Across this resistance there exist a potential difference, called "the lost volts" (as shall be seen in the example. The "full" main current "moves through the battery".

2.1 Find the resistance of the unknown resistor X in the circuit if the main current in the circuit is equal to 1,6A.
(Background knowledge: Two voltmeter readings for the battery are given, namely: E and V.
E = Emf (called the electro-motor force) = 9 volt (V); and
V = Potential difference across the external circuits' components = 8,4V.
V_1 = "Lost volts" = E – V = 9,0 - 8,4 = 0,6V - this potential difference is across the internal resistance of the battery r)

The total resistance of the external circuit can easily be calculated because the pd across the external circuit is known as well as the total current.

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Therefore:

$$R = \frac{V}{I} = \frac{8,4}{1,6} = 5,25 \text{ ohm } (\Omega)$$

Remember: totoal external pd and main current is used find the total resistance of the external circuit

Now: Find Rp as follows:

$$\frac{1}{Rp} = \frac{1}{R1} + \frac{1}{R2}$$

Note: Two branches

$$= \frac{1}{6} + \frac{1}{(2 + 6)}$$

Two resistors in one of the branches

$$= \frac{4 + 3}{24}$$

$$= \frac{7}{24}$$

$$Rp = \frac{24}{7} = 3,43 \text{ ohm}$$

We now know the total external resistance (5,25 ohm) and the Rp of the parallel resistors. The unknown value is calculated as follows:

$Rt = Rp + Rx$

$Rx = Rt - Rx = 5,25 - 3,43 = 1,82$ ohm

2.2 Find the value o r (r is the internal resistance of the battery)

$$r = \frac{V - E}{I}$$

$$r = \frac{0,6}{1,6} \quad 0,375 \text{ ohm}$$

Remember V-E is the so called "lost volts" because it can only be determined "indirectly" and not measured by a voltmeter

1.3 Determine the readings on V3 and V4 ("the voltage between lines").

First: We need I_1 and I_2.

Pd across Rp: $V = Im \times Rp = 1.6 \times 3,43 = 5,49V$

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$$I_1 = \frac{V_{rp}}{R_1} \quad \text{and} \quad I_2 = \frac{V_{rp}}{R_2}$$

$$= \frac{5,49}{8} \qquad = \frac{5,49}{6}$$

$$= 0,68A \qquad = 0,92A$$

Secondly; We require the pd's (or voltmeter readings across the components):

$V_1 = I_1 \times R = 0,68 \times 2 = 1,36V$
$V_2 = I_1 \times R = 0,68 \times 6 = 4.08V$
$V_5 = I_2 \times R = 0,92 \times 6 = 5,52V$

$V_3 = V_5 - V_1 = 5,52 - 1,36 = 4.16V$
$V_4 = V_2 - 0 = 4.08V$

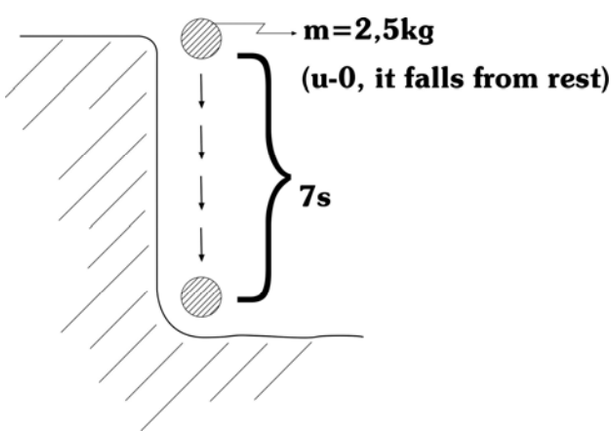
Note: Between the lines the potential difference is the same (see Example)

FREE FALLING OBJECTS AND PROJECTED MOTION

INTRODUCTION:

"Free fall" means that the object is allowed to fall in the earths g-field freely, for example:

1. An object falls from rest from a cliff and reaches the ground 7 seconds later.



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1.1 Determine the velocity when the object reaches the ground.

(Formulas: $v = u + gt$; $v^2 = u^2 + 2gs$ and $s = ut + \frac{1}{2} gt^2$

$v = u + gt = 0 + (10 \times 7) = 70\ m.s^{-1}$

1.2 What is the height of the cliff?

$s = ut + \frac{1}{2} gt^2 = (0 \times 7) + (0,5 \times 10 \times 7^2)$

$= 0 + 245 = 245\ m$

1.3 Determine the kinetic energy (E_k) of the object just before it reaches the ground if the objects' mass is equal to 2,5 kg.

$E_k = \frac{1}{2} mv^2 = 0,5 \times 2,5 \times 70^2 = 6125$

Note: Kinetic energy is energy of motion of objects

1.4 Determine the potential energy (E_p) of the object when it is stationary on top of the cliff.

$E_p = mgh = 2,5 \times 10 \times 245 = 6125$ joule (J)

(Note: Potential energy is energy stored in the objects' position in the g-field)

1.5 Notes:

- The E_k the moment before the object hits the ground is equal to the E_p stored in the objects position on the cliff.
- The total mechanical energy at any point in a closed system = $E_k + E_p$.

1.6 Determine the kinetic energy of the object when it is 45m above the ground.

First: $E_p = mgh = 2,5 \times 10 \times 45 = 1125J$ (Remember at 45m above the ground).

But we are required to find E_k.

Secondly: Total ME = $E_k + E_p$ (at any point in a closed system)

$6125 = E_k + 1125$; $E_k = 6125 - 1125 = 5000J$

A check on our calculation:

$v^2 = u^2 + 2gs$

$V^2 = 0 + (2 \times 10 \times 200)$ (The distance fallen to reach 45m is 200m)

$= 4000$

$v = 63,25\ m.s^{-1}$ (Velocity 45m above the ground)

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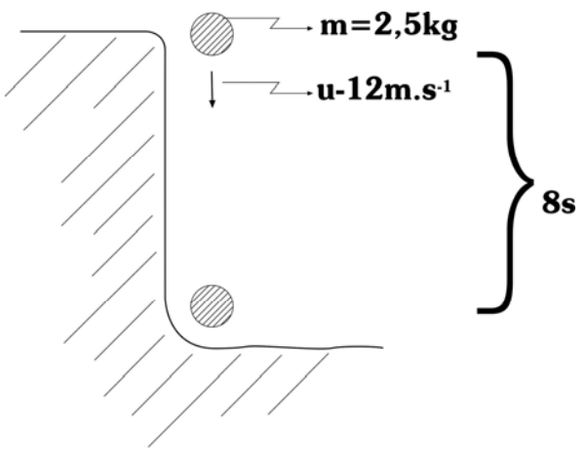
$$E_k = \frac{1}{2} mv^2$$

$$= 0,5 \times 2,5 \times 63,25^2$$

$$= 5001J$$ (Kinetic energy of the object 45m above the ground)

"Projected motion" means that the object is given an initial velocity.

2. An object of mass 2,3 kg is projected vertically downward with an initial velocity of 12 $m.s^{-1}$.



2.1 Find the object's velocity just before it reaches the ground if it takes 8s to fall to the ground.

$v = u + gt = 12 + (10 \times 8) = 12 + 80 = 92\ m.s^{-1}$.

2.2 Find the height of the cliff.

$$s = ut + \frac{1}{2} gt^2 = (12 \times 8) + (0,5 \times 10 \times 8^2) = 96 + 320 = 416m$$

(Remember: u is always initial velocity and v is always the final velocity).

3. An air balloon moves upward at a constant velocity of 2,5 $m.s^{-1}$. A student projects an object vertically upward from the air balloon with a constant velocity of 1,6 $m.s^{-1}$.

3.1 Velocity of the object relative to the ground.

$v = v_1 + v_2 = 2,5 + 1,6 = 4,1\ m.s^{-1}$

(Both the velocities are upward; therefore the relative velocity upward to the ground is 4,1 $m.s^{-1}$.)

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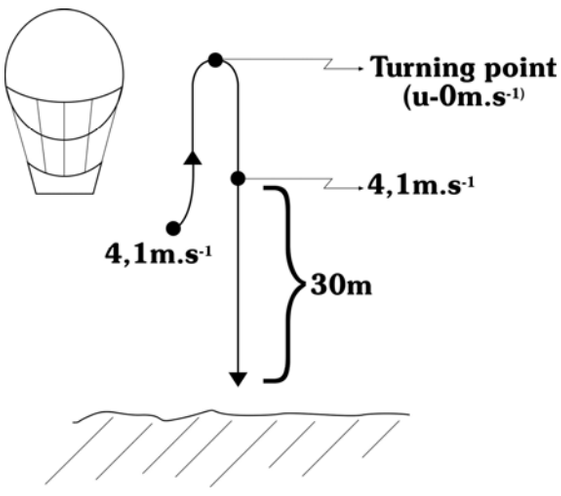
Now let us make a sketch first:



3.2 Find the velocity just before the object hits the ground.

$v^2 = u^2 + 2gs$ (See sketch for initial velocity and height)

$= 4,1^2 + (2 \times 10 \times 30)$

$= 16,81 + 600$

$= 616,81$

$v = 24,84 \ m.s^{-1}$

3.3 Find the total time for the object being in the air before it hits the ground.

First find the time for the total downward fall from the turning point as follows:

$v = u + gt$

$= 0 + gt$

$v = gt$

$gt = v$

$$t = \frac{v}{g}$$

$$= \frac{24,84}{10}$$

$= 2,48s$

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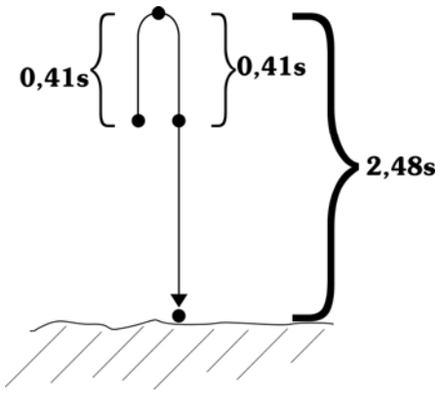


But the object moved upward before its downward fall:

$$v = u + gt$$

$$v - u = gt$$

$$gt = v - u$$

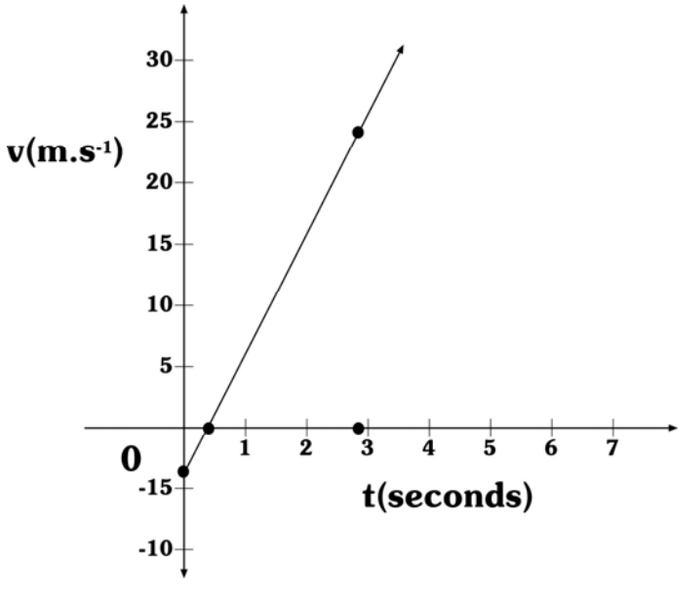
$$t = \frac{v - u}{g}$$

$$= \frac{4,1 - 0}{10}$$

$$= 0,41s$$

Total time in the air = 2,48 + 0,41 = 2,49s (See sketch for explanation)

3.4 Draw a v/t-graph for the motion of the object:



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Remember: The negative part (4,1 $m.s^{-1}$) represents the upard motion and the rest (+) the fall forwards the earth.

MOMENTUM OF OBJECTS

Theory:

- The momentum of an object is defined as the property an object has due to the product of its mass and its velocity.
- Basic formulas: $p = mv$ (Note: Because velocity is part of the formula and it possesses direction , momentum also is a vector quantity).
- When an objects velocity changes, the momentum also changes:

$\Delta p = mv - mu$ (Δp means "change" ; therefore change in momentum; final momentum – initial momentum).

$\Delta p = F\Delta t$ = Impuls (where Δt = change in time or elapsed time). Impuls can be regarded as "the magnitude or impact or effect of the collision" of the object. Remember it can be defined as the product of the F applied and the time of contact during the collision.

1. A ball of mass 210g moves forward (Note: Taken as positive) at 14 $m.s^{-1}$ towards a wall, hits it and return (Note: Taken as negative) at 12 $m.s^{-1}$.

1.1 Find the momentum of the object when it approaches the wall.

$$p = mv = \frac{210}{1000} \times 14 = 0{,}21 \times 14 = 2{,}94 \text{ kg.m.s}^{-1}$$

Note: Momentum has direction, therefore it is a vector. Remember: We devide the mass (given by g)by 1000 to convert it to kg. The Unit need not to be remembered; it can be deducted from the given formula p=mv

1.2 Find the change in momentum during the collision.

$\Delta p = mv - mu = (0{,}21 \times [-12]) - (0{,}21 \times 14) = -2{,}52 - 2{,}94 = -5{,}46$ kg.m.s$^-$ (left)

Remember: Negative indicates to the left.

1.3 HG ONLY: Determine the impuls.

Impuls = $\Delta p = F\Delta t = 5{,}46$N.s (left) (Note: The negative sign indicates left).

Remember: The unit N.s need not to be remembered: See formula)We also expect the force and impuls to be left – the ball returns to us (the throwers of the ball).

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1.4 HG ONLY: The collision with the wall takes 0,015 seconds. Find the force of the wall on the object.

Impuls = $F\Delta t = \Delta p$ (given)

$F\Delta t = \Delta p$ (Swop the formula around)

$$F = \frac{\Delta p}{\Delta t}$$

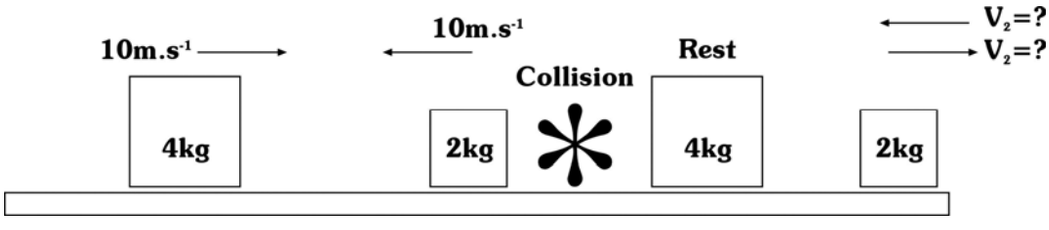
$$= \frac{-5,46}{0,015}$$

$= -364N$ (left: see negative sign)

Theory about the conservation of momentum:

- In a closed system, the total linear momentum is conserved (meaning that no "momentum" is "lost".)
- In a closed system, the Momentum before the collision = the momentum after the collision.
- If a sketch of the colliding bodies is not made, the problems can be very confusing.

1.2 Two objects with mass 4 kg and 2 kg moves in opposite directions at 10 $m.s^{-1}$ when they collide elastically (they do not "stick" together). The 4 kg comes to a standstill. In what direction and at what velocity is the other one moving?



Momentum before the collision = Momentum after the collision

$(4 \times 10) + (2 \times [-10]) = (4 \times 0) + (2 \times v_2)$

$40 - 20 = 0 + 2v_2$

$20 = 2v_2$

$2v_2 = 20$

$$v_2 = \frac{20}{10}$$

$v_2 = 2\ m.s^{-1}$

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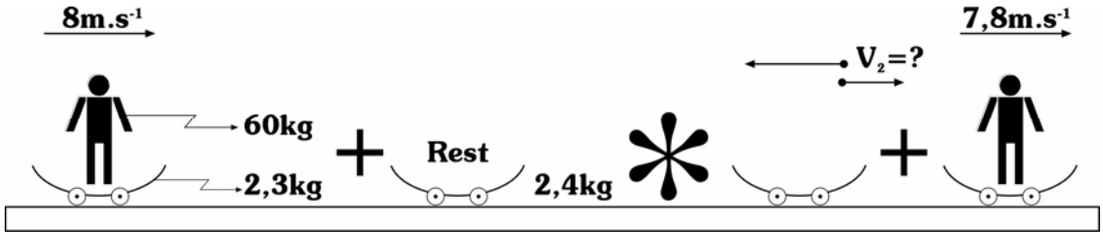
The object moves to the right (remember "to the right" was taken as positive; therefore the positive answer means "to the right") at $2 m.s^{-1}$.

1.3 A skate board rider of mass 60 kg skates forward with a skate board of mass 2,3 kg with a velocity of 8 $m.s^{-1}$ and then jumps unto a stationary skate board of mass 2,4 kg. His velocity on the new skate board after the jump is 7,8 $m.s^{-1}$. Find the velocity of the first skate board.



Momentum before the collision = Momentum after the collision

$([2,3 + 60] \times 8)+(2,4 \times 0)= [2,4 + 60] \times 7.8) + 2,3v_2$

$498,4 + 0 = 486,72 + 2,3v_2$

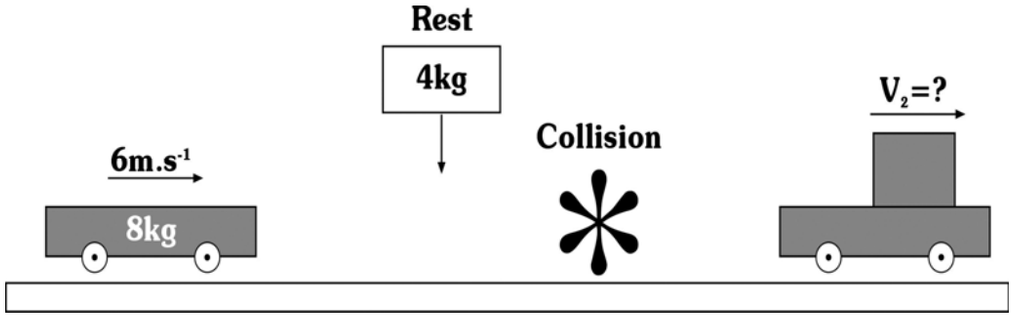
$2,3v_2 = 498,4 - 486,72$

$2,3v_2 = 11,6$

$$v_2 = \frac{11,6}{2,3}$$

$= 5,1 m.s^{-1}$ (to the right or forward)

1.4 A trolley of mass 8 kg moves forward with a velocity of 6 $m.s^{-1}$ when an object of mass 4 kg is dropped onto it. What is the velocity of the combination? (Remember: Now the collision is non-elastic).



Momentum before the collision = Momentum after the collision

$(8 \times 6) + (4 \times 0) = ([8 + 4] \times v_2)$

$12v_2 = 48$

$v_2 = 4 m.s^{-1}$

Remember: The negative part ($4,1 m.s^{-1}$) represents the upward motion and the rest (+) the fall towards the earth.

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