

## Dunearn Secondary School

### Pure Chemistry Prelim Exam Paper 1-2011

Q1: Student 4 had more base than the other due to.....

- A) Acid in burette, so washing it in acid does not matter..
- B) Base was in Conical flask, so washing it with base possibly increase the amount of base which needed more acid to neutralize.
- C) Pipette used to measure base was washed out with base, it does not matter
- D) Reading from top of meniscus should not matter so long as it read out is consistent before and after the titration

→ Option **B** is correct—Ans

Q2: Purity should be tested quantitatively.

Taste is qualitative, Acidity depends on the dissociation constant of acid, rate of reaction cannot provide quantitative measure of impurities. Only BP is good quantitative indication of purity of ethanoic acid. Presence of dissolved salts increases the BP. If there is mixture of different liquid phase, there would be multiple BP.

→ Option **C** is correct—Ans

Q3: Compute R<sub>f</sub> for the 4 dots

0.9, 0.8, 0.4, 0.3 → K101, K151

→ Option **A** is correct—Ans

Q4:

To separate immiscible liquid phases, use separating funnel.

To separate KCl from water, use evaporation to dryness since KCl is very stable salts that not easily decompose on heat.

→ Option **D** is correct—Ans

Q5:

A; likely to be hydrocarbon → compound

B; likely to be metal → element

C; likely to be carbonate → Compound

D; a mixture of compound or elements

→ Option **B** is element—Ans

Q6:

Cooling curve means heat released to surrounding,  $t_2$ - $t_3$ , progressively formation of solid.....the curve should cover from liquid phase to solid phase.. (Not gas to liquid phase was shown)

→ Option **C** is correct—Ans

Q7:

MP - BP

-170 -108

-150 Liquid phase

-100 gas phase

→ Option **A** is correct—Ans

Q8: Isotope same p, but different n (e might be difference due to ions)

G=H

→ Option **B** is correct—Ans

Q9:

K L M N

Group 7 8 1 2

Possible ionic salts=MK, NK<sub>2</sub>

→ Option **C** is correct—Ans

Q10: use electronic configuration

P Q R S

2,3 2,8,2 2,8,5 2,8,8

Q is metallic which lose e<sup>-</sup> more easily than R from group 5

→ Option **B** is correct—Ans

Q11:

T: Ammonia soluble in water to form  $\text{NH}_4^+\text{Cl}^-$

U: Graphite which is conductive carbon insoluble

V: Metal conduct both in solid and liquid phase

W: Salt conduct only in molten and aqueous

→ Option **A** is correct—Ans

Q12:

0.8 g of  $\text{O}_2 = 0.8/32 \text{ mol} = 0.8 * 24/32 \text{ dm}^3$

gas X in same volume =  $2.2 \text{ g} = 2.2/\text{mr} \text{ mol} = 24 * 2.2/\text{mr}$

$0.8/32 = 2.2/\text{mr} \rightarrow \text{mr} = 88$

→ Option **D** is correct—Ans

Q13:

Mole of  $\text{ZnCl}_2 = 23 / (65 + 35.5 * 2) = 23 / 136 = \text{mole of ZnO}$

Actual Mass of  $\text{ZnO} = (65 + 16) * 23 / 136 \text{ g} = 13.7 \text{ g}$

$\% \text{Zn} = 13.7 / 19 = 72.1\%$

→ Option **C** is correct—Ans

Q14: All gaseous reactant so go by volume ratio

Oxygen in excess (  $50 - 30 = 20 \text{ cc}$  ),  $\text{CO}_2$  formed 10 cc,  $\text{SO}_2$  formed 20 cc

Total products + excess O<sub>2</sub> = 20 = 10 + 20 = 50 cc

Treatment of NaOH removed all CO<sub>2</sub> + SO<sub>2</sub> so remaining only 20CC oxygen

→ Option **C** is correct—Ans

Q15:

Common yellow pH range

3.5 to 6

5.5 to 6

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5.5 to 6 --solution F

→ Option **D** is correct—Ans

Q16:

Neutralization so solid added should be a base:

Na<sub>2</sub>SO<sub>4</sub> not a base

K<sub>2</sub>O + HCl → 2KCl + H<sub>2</sub>O in excess will be alkaline due to K<sub>2</sub>O soluble

CaCO<sub>3</sub> + 2HCl → CaCl<sub>2</sub> + CO<sub>2</sub> + H<sub>2</sub>O excess insoluble thus neutral

ZnO + 2HCl → ZnCl<sub>2</sub> + H<sub>2</sub>O excess insoluble thus neutral

→ Option **B** is correct—Ans

Q17:

PbSO<sub>4</sub> insoluble, use ppt method. Noted that use solid + acid would not work as the PbSO<sub>4</sub> coating formed will cease the reaction.

→ Option **D** is correct—Ans

Q18:

H & J Alkaline due to reaction with ammonia salts, J strong base highest pH due to high conductivity

I & K acid, I-strong acid lowest pH

pH value I < K < H < J

→ Option **A** is correct—Ans

Q19:

Presence of NO<sub>3</sub> ions by adding NaOH + Aluminum, without Al, no NH<sub>3</sub> so the salt should not be ammonia salt

→ Option **A** is correct—Ans

Q20:

P less reactive than M; N more reactive than P, M more reactive than N

Reactivity M > N > P

→ Option **D** is correct—Ans

Q21:

P:Raw ore + CaCO<sub>3</sub> + Coke    Q:waste gases    R:Hot air    S: Slag    T=iron

→ Option **B** is correct—Ans

Q22:

Group 1 metal : they has different electrons config but same valency, they have different reactivity and physical properties. But they form same type of bond with hydrogen.

→ Option **A** is correct—Ans

### **The Group 1 hydrides**

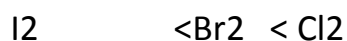
#### **Saline (salt-like) hydrides**

The hydrides of Group 1 metals are white crystalline solids which contain the metal ions and hydride ions,  $H^-$ . They have exactly the same crystal structure as sodium chloride - that's why they are called saline or salt-like hydrides.

Because they can react violently with water or moist air, they are normally supplied as suspensions in mineral oil.

Q23: Displacement reaction of halogen

W2 least reactive <U2 <V2 most reactive displace the others



Option **A** is correct--Ans

Q24: Order of element in periodic table

Metallic (form basic oxide) → amphoteric → Nonmetal (form acidic oxides)

Y

Z

X

→ Option **D** is correct—Ans

Q25: Redox

$2\text{NH}_4\text{OH} + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4 + \text{H}_2\text{O}$  neutralization reaction

$\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$  Fe was reduced Redox

$\text{AgNO}_3 + \text{HCl} \rightarrow \text{AgCl ppt} + \text{HNO}_3$  Precipitation reaction

$\text{SO}_2 + \text{Ca}(\text{OH})_2 \rightarrow \text{CaSO}_3 + \text{H}_2\text{O}$  Acid/base reaction

→ Option **B** is correct—Ans

Q26: Redox

A) Hydrogen OS from 0 to +1 Oxidized → RA

B) Fluorine OS from 0 to -1 Reduced → OA

C) Cu OS from +2 to 0 Reduced → OA

D) Mn OS from +4 to +2 Reduced → OA

→ Option **A** is correct—Ans

Q27:

Nitrogen OS from +2 to 0, Carbon OS from +2 to +4 → Redox

Typical catalytic converter



→ Option **D** is correct—Ans

Q28:

CO<sub>2</sub> : greenhouse gas, acidic

CH<sub>4</sub>: greenhouse gas

SO<sub>2</sub>: Acid rain

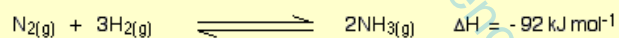
→ Option **D** is correct—Ans

Q29: Haber process

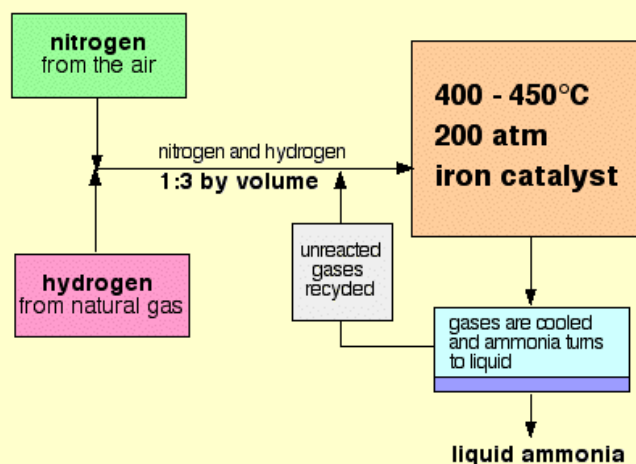
$\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$  250 atm, 450 C, iron catalyst

→ Option **A** is correct—Ans

The Haber Process combines nitrogen from the air with hydrogen derived mainly from natural gas (methane) into ammonia. The reaction is reversible and the production of ammonia is exothermic.



A flow scheme for the Haber Process looks like this:



Q30: CuSO<sub>4</sub> electrolysis

Cathode(-): Cu<sup>2+</sup> deposited regardless of electrode material

Anode (+): 4OH<sup>-</sup> → 2H<sub>2</sub>O + O<sub>2</sub> + 4e for inert electrode, Cu → Cu<sup>2+</sup> + 2e for copper electrode

→ Option **D** is correct—Ans

Q31:

A) 4x3=12 mol of e

B) 5x1=5 mol of e

C) 6x2=12 mol of e

D) 7x2=14 mol of e

→ Option **D** is correct—Ans

Q32:

I) No bond breaking, only bond forming → Exo

II) CO<sub>2</sub> + H<sub>2</sub>O photosynthesis → Endo

III) Neutralization → Exo

→ Option **B** is correct—Ans

Q33: Low bond energy → weak bond

A) Si-Si is most readily broken

B) Si-Si had lower total energy → More stable

C) 1 mol of graphite had more than 1 mol of C-C bond as it is giant molecule

D) CH<sub>4</sub> had 4 C-H (total energy=4\*413) , SiH<sub>4</sub> had 4 Si-H ( 4\*318), lower total energy is more stable.

→ Option B is correct—Ans

Q34: based on the KE model

### Increasing the collision frequency

Particles can only react when they collide. If you heat a substance, the particles move faster and so collide more frequently. That will speed up the rate of reaction.

That seems a fairly straightforward explanation until you look at the numbers!

It turns out that the frequency of two-particle collisions in gases is proportional to the square root of the kelvin temperature. If you increase the temperature from 293 K to 303 K (20°C to 30°C), you will increase the collision frequency by a factor of:

$$\sqrt{\frac{303}{293}} = 1.017$$

That's an increase of 1.7% for a 10° rise. The rate of reaction will probably have doubled for that increase in temperature - in other words, an increase of about 100%. The effect of increasing collision frequency on the rate of the reaction is very minor. The important effect is quite different . . .

As you increase the temperature the rate of reaction increases. As a rough approximation, for many reactions happening at around room temperature, the rate of reaction doubles for every 10°C rise in temperature.

You have to be careful not to take this too literally. It doesn't apply to all reactions. Even where it is approximately true, it may be that the rate doubles every 9°C or 11°C or whatever. The number of degrees needed to double the rate will also change gradually as the temperature increases.

→ Option **A** is correct--Ans

Q35:

I) False: Not all collisions result in reaction

II) True: Higher concentration increase SOR

III) False: For gases reactant, higher pressure is equivalent to higher conc.

IV) True: effect of Temperature on SOR is typically exponential ( conservatively double for every 10 degree K increase in temp)

→ Option **C** is correct—Ans

Q36: assume same amount of Zn was used

A) 0.08 g/cm<sup>3</sup> B) 0.04 g/cm<sup>3</sup> C) 0.2 g/cm<sup>3</sup> D) 0.16 g/cm<sup>3</sup>

→ Option **C** is correct--Ans

Q37: Fractional distillation, higher BP temp range → heavier molecule + higher viscosity → lower flammability

A) False

B) True: smaller molecule burn more readily

C) False

D) False

→ Option **B** is correct—Ans

Q38: Cracking  $C_{10}H_{22} \rightarrow 4C_2H_4 + C_2H_6$

→ Option **B** is correct—Ans

Q39:

-OH & C=C: alcohol & alkene functional group

→ option **C** is correct—Ans

Q40: formation of ester by reaction of alcohol + Carboxylic acid

A: Ester

B: Ester

C: Non-ester

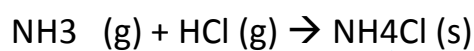
D: Ester

→ Option C is correct –Ans

## Dunearn Secondary School

### Pure Chemistry Prelim Exam Paper 3-2011

A1:



Mr=17      36.5

Heavier HCl molecules diffuse slower than NH<sub>3</sub> thus NH<sub>4</sub>Cl form nearer to HCl end.

$$3.65/36.5=0.01 \text{ mol HCl}$$

3.4/17= 0.2 mol NH<sub>3</sub> (excess) thus litmus paper red to blue due to alkaline NH<sub>4</sub>OH.

A2:

a) OS of copper

b) 2+ in CuSO<sub>4</sub>, 0 in Cu, 1+ in Cu<sub>2</sub>O

Cu(I) was both reduced to Cu metal and oxidized to Cu(II) → redox

c) Test for presence of OA using KI solution, change from colorless to brown in the presence of OA.

A3:

a)

i) Enthalpy change ii) 2O<sub>3</sub> iii) Activation energy iv) 3O<sub>2</sub>

b) CFC are the source of chlorine in the ozone layer

The overall amount of ozone in the stratosphere is determined by a balance between photochemical production and recombination.

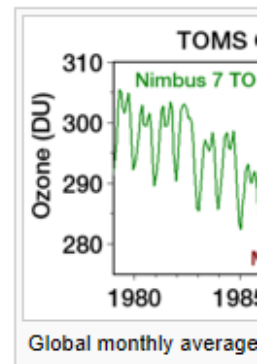
Ozone can be destroyed by a number of **free radical** catalysts, the most important of which are the **hydroxyl radical** (OH·), the **nitric oxide radical** (NO·), the atomic **chlorine** ion (Cl·) and the atomic **bromine** ion (Br·). All of these have both natural and man-made sources; at the present time, most of the OH· and NO· in the stratosphere is of natural origin, but human activity has dramatically increased the levels of chlorine and bromine. These elements are found in certain stable organic compounds, especially **chlorofluorocarbons** (CFCs), which may find their way to the **stratosphere** without being destroyed in the troposphere due to their low reactivity. Once in the stratosphere, the Cl and Br atoms are liberated from the parent compounds by the action of ultraviolet light, e.g.



The Cl and Br atoms can then destroy ozone molecules through a variety of **catalytic** cycles. In the simplest example of such a cycle,<sup>[4]</sup> a chlorine atom reacts with an ozone molecule, taking an oxygen atom with it (forming ClO) and leaving a normal oxygen molecule. The chlorine monoxide (i.e., the ClO) can react with a second molecule of ozone (i.e., O<sub>3</sub>) to yield another chlorine atom and two molecules of oxygen. The chemical shorthand for these gas-phase reactions is:

- $\text{Cl} + \text{O}_3 \rightarrow \text{ClO} + \text{O}_2$  – The chlorine atom changes an ozone molecule to ordinary oxygen
- $\text{ClO} + \text{O}_3 \rightarrow \text{Cl} + 2 \text{O}_2$  – The ClO from the previous reaction destroys a second ozone molecule and recreates the original chlorine atom, which can repeat the first reaction and continue to destroy ozone

The overall effect is a decrease in the amount of ozone. More complicated mechanisms have been discovered that lead to ozone destruction in the lower stratosphere as well.



**Chlorofluorocarbons** (CFCs) and other halogenated ozone depleting substances (ODS) are mainly responsible for man-made chemical ozone depletion. The total amount of effective **halogens** (**chlorine** and **bromine**) in the stratosphere can be calculated and are known as the **equivalent effective stratospheric chlorine** (EESC).<sup>[12]</sup>

Hole in ozone layer allow harmful UV light from sun to enter which can cause cancerous cells.

A4:

a)  $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$  iron easily oxidize when heated in air

b)  $3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$  by volume calculation since all gaseous

75cc    25cc    50cc

$\text{NH}_3$  produced =  $50 \times 15/100 = 7.5$  CC, remaining 63.75cc of  $\text{H}_2$ , 21.25 cc of  $\text{N}_2$

## Le Chateliers' principle

state that if pressure was increase, the equilibrium point shift to the right, if temperature was increase, the equilibrium point shift to the left for exothermic reaction, thus an optimum setting of temperature, pressure and catalyst is required to maximum yield.

### Reaction rate and equilibrium

[edit]

There are two opposing considerations in this synthesis: the position of the equilibrium and the **rate of reaction**. At room temperature, the reaction is slow and the obvious solution is to raise the temperature. This may increase the rate of the reaction but, since the reaction is **exothermic**, it also has the effect, according to **Le Chatelier's principle**, of favouring the reverse reaction and thus reducing the amount of product, given by:

$$K_{\text{eq}} = \frac{P_{\text{NH}_3}^2}{P_{\text{N}_2} \cdot P_{\text{H}_2}^3}$$

As the temperature increases, the **equilibrium** is shifted and hence, the amount of product drops dramatically according to the **Van't Hoff equation**. Thus one might suppose that a low temperature is to be used and some other means to increase rate. However, the catalyst itself requires a temperature of at least 400 °C to be efficient.

**Pressure** is the obvious choice to favour the forward reaction because there are 4 moles of reactant for every 2 moles of product (see **entropy**), and the pressure used (around 200 atm) alters the equilibrium concentrations to give a profitable yield.

Economically, though, pressure is an expensive commodity. Pipes and reaction vessels need to be strengthened, valves more rigorous, and there are safety considerations of working at 200 atm. In addition, running pumps and compressors takes considerable energy. Thus the compromise used gives a single pass yield of around 15%.

Another way to increase the yield of the reaction would be to remove the product (i.e. ammonia gas) from the system. In practice, gaseous ammonia is not removed from the reactor itself, since the temperature is too high, but it is removed from the equilibrium mixture of gases leaving the reaction vessel. The hot gases are cooled enough, whilst maintaining a high pressure, for the ammonia to condense and be removed as liquid. Unreacted hydrogen and nitrogen gases are then returned to the reaction vessel to undergo further reaction.

Variation in  $K_{\text{eq}}$  for the equilibrium  
 $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$   
as a function of temperature  
[10]

Temperature (°C)	$K_{\text{eq}}$
300	$4.34 \times 10^{-3}$
400	$1.64 \times 10^{-4}$
450	$4.51 \times 10^{-5}$
500	$1.45 \times 10^{-5}$
550	$5.38 \times 10^{-6}$
600	$2.25 \times 10^{-6}$

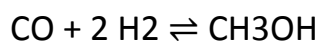
In the lab, the pressure was lower than industrial haber process, thus the amount of  $\text{NH}_3$  produce would be lesser than the computed value above.



## **Effect of change in concentration**

Changing the concentration of an ingredient will shift the equilibrium to the side that would reduce that change in concentration. The chemical system will attempt to partially oppose the change affected to the original state of equilibrium. In turn, the rate of reaction, extent and yield of products will be altered corresponding to the impact on the system.

This can be illustrated by the equilibrium of carbon monoxide and hydrogen gas, reacting to form methanol.

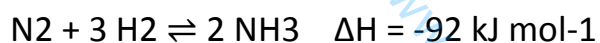


Suppose we were to increase the concentration of CO in the system. Using Le Chatelier's principle, we can predict that the amount of methanol will increase, decreasing the total change in CO. If we are to add a species to the overall reaction, the reaction will favor the side opposing the addition of the species. Likewise, the subtraction of a species would cause the reaction to fill the "gap" and favor the side where the species was reduced. This observation is supported by the collision theory. As the concentration of CO is increased, the frequency of successful collisions of that reactant would increase also, allowing for an increase in forward reaction, and generation of the product. Even if a desired product is not thermodynamically favored, the end-product can be obtained if it is continuously removed from the solution.

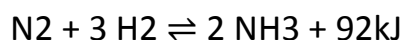
## **Effect of change in temperature**

The effect of changing the temperature in the equilibrium can be made clear by incorporating heat as either a reactant or a product. When the reaction is exothermic ( $\Delta H$  is negative, puts energy out), we include heat as a product, and, when the reaction is endothermic ( $\Delta H$  is positive, takes energy in), we include it as a reactant. Hence, we can determine whether increasing or decreasing the temperature would favour the forward or reverse reaction by applying the same principle as with concentration changes.

For example, the reaction of nitrogen gas with hydrogen gas. This is a reversible reaction, in which the two gases react to form ammonia:



If you put heat as a product:



This is an exothermic reaction (hence the minus sign) when producing ammonia. If we were to lower the temperature, the equilibrium would shift to produce more heat. Since making ammonia is exothermic, this would favour the production of more ammonia. In practice, in the Haber process, the temperature is set at a compromise value, so ammonia is made quickly, even though less would be present at equilibrium.

In exothermic reactions, increase in temperature decreases the equilibrium constant,  $K$ , whereas, in endothermic reactions, increase in temperature increases the  $K$  value.

The value of  $K$  changes with temperature. In the endothermic reaction  $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ , the equilibrium position can be shifted by changing the temperature. When heat is added and the temperature increases, the reaction shifts to the right and the flask turns reddish brown due to an increase in  $\text{NO}_2$ . When heat is

removed and the temperature decreases, the reaction shifts to the left and flask turns colorless due to an increase in  $\text{N}_2\text{O}_4$ . This demonstrates Le Chatelier's Principle because the equilibrium shifts in the direction that consumes energy.

### **Effect of change in pressure**

Changes in pressure are attributable to changes in volume. The equilibrium concentrations of the products and reactants do not directly depend on the pressure subjected to the system. However, a change in pressure due to a change in volume of the system will shift the equilibrium.

Considering the reaction of nitrogen gas with hydrogen gas to form ammonia:



Note the number of moles of gas on the left-hand side and the number of moles of gas on the right-hand side. When the volume of the system is changed, the partial pressures of the gases change. If we were to decrease pressure by increasing volume, the equilibrium of the above reaction will shift to the left, because the reactant side has greater number of moles than does the product side. The system tries to counteract the decrease in partial pressure of gas molecules by shifting to the side that exerts greater pressure. Similarly, if we were to increase pressure by decreasing volume, the equilibrium shifts to the right, counteracting the pressure increase by shifting to the side with fewer moles of gas that exert less pressure. If the volume is increased because there are more moles of gas on the reactant side, this change is more significant in the denominator of the equilibrium constant expression, causing a shift in equilibrium.

Thus, an increase in system pressure due to decreasing volume causes the reaction to shift to the side with the fewer moles of gas.[3] A decrease in pressure due to increasing volume causes the reaction to shift to the side with more moles of gas. There is no effect on a reaction where the number of moles of gas is the same on each side of the chemical equation.

### **Effect of adding an inert gas**

An inert gas (or noble gas) such as helium is one that does not react with other elements or compounds. Adding an inert gas into a gas-phase equilibrium at constant volume does not result in a shift.[4] This is because the addition of a non-reactive gas does not change the partial pressures of the other gases in the container. While it is true that the total pressure of the system increases, the total pressure does not have any effect on the equilibrium constant; rather, it is a change in partial pressures that will cause a shift in the equilibrium. If, however, the volume is allowed to increase in the process, the partial pressures of all gases would be decreased resulting in a shift towards the side with the greater number of moles of gas. There is a short form to remember this: LBMF (little boy married fiona); L stands for less pressure, B - backward reaction, M - more pressure, and F - forward reaction.

### **Effect of a catalyst**

A catalyst speeds up the rate of a reaction by providing additional mechanism(s). Adding a catalyst allows for alternative pathways to be made, where the particles can be adsorbed onto the catalyst temporarily before being re-bonded into a new arrangement. The intended effect in adding a catalyst is to lower the activation energy, which frequently increases the rate of reaction. However, the activation energy is lowered by the same amount for the forward and reverse reactions. There is the same increase in reaction rates for both reactions. As a result, a catalyst does not affect the position of the equilibrium. It only affects the time or energy that is required to achieve equilibrium.

A5: Careful 4NaOH had residue 2Na<sub>2</sub>O when strong heated to dryness

Test5-Q contain NO<sub>3</sub>-

Test 4- ZnO (W) is yellow when hot but white when cooled, the precipitation was ZnCO<sub>3</sub> (V) insoluble in aqueous solution.

Test 3- BaSO<sub>4</sub> (U) insoluble so P contain SO<sub>4</sub> ions and Fe(II) ions, FeSO<sub>4</sub> (P)

Test 2- NH<sub>4</sub>OH + Zn(NO<sub>3</sub>)<sub>2</sub> (Q) → Zn(OH)<sub>2</sub> + 2NH<sub>4</sub>NO<sub>3</sub>

Test 1- NH<sub>4</sub>OH (R) evaporate completely without residue

A6:

a) Mg + 2HCl → MgCl<sub>2</sub> + H<sub>2</sub>

EXP(I) mol mg=4.8/24=0.2 mol mol HCl=0.1\*1.5=0.15 mol → HCl is limiting

Mol H<sub>2</sub> produced=0.15/2=0.075

b) H<sub>2</sub>SO<sub>4</sub> produce 2H<sup>+</sup> ion per molecule compared to HCl only 1H<sup>+</sup> per molecule, thus H<sub>2</sub>SO<sub>4</sub> had higher concentration of H<sup>+</sup> thus more H<sub>2</sub> produced as well as faster rate of reaction. The final volume of H<sub>2</sub> produced will be more.

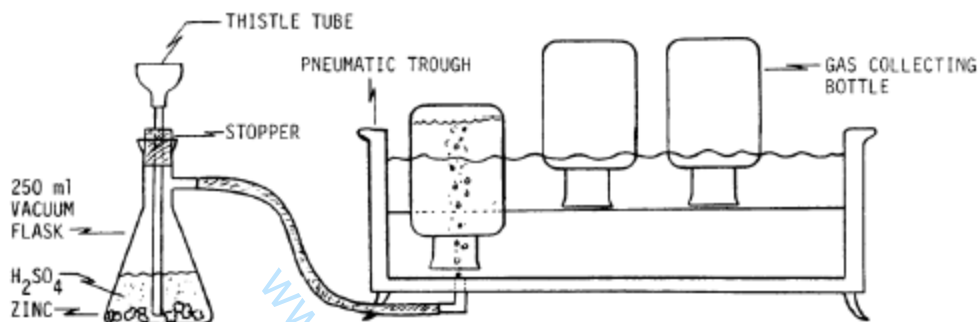
Mg + H<sub>2</sub>SO<sub>4</sub> → MgSO<sub>4</sub> + H<sub>2</sub>

EXP(I) mol mg=4.8/24=0.2 mol mol H<sub>2</sub>SO<sub>4</sub>=0.1\*1.5=0.15 mol → H<sub>2</sub>SO<sub>4</sub> is limiting

Mol H<sub>2</sub> produced=0.15

c) EXP(II)  $\text{CuSO}_4 + \text{Mg} \rightarrow \text{MgSO}_4 + \text{Cu}$  due to metal displacement reaction

The presence of  $\text{Cu}^{2+}$  act as catalyst increasing the formation of  $\text{H}_2$ . Note that Cu is a transition metal and it can exist as Cu(I) or Cu(II) which it make possible to play the role of catalyst.



2. Add approximately eight grams of mossy zinc (Zn) to the bottom of the 250 ml vacuum flask.
3. Make all connections airtight to prevent leakage of the hydrogen gas ( $\text{H}_2$ ).
4. Push the thistle tube to within approximately 0.5cm (1/4") of the bottom of the flask.
5. Add 2 ml of 0.1M copper sulfate ( $\text{CuSO}_4$ ) to the thistle tube. The copper sulfate acts as a catalyst in the production of hydrogen gas (speeds up rate of production).
6. Fill a gas-collecting bottle with water and stopper it. Invert the bottle (stopper and downward), and place in the water-filled pneumatic trough. Remove the stopper. The water should remain in the bottle as long as the neck of the bottle remains underneath the surface of the water.

A7:

a) cell Q:  $\text{Ag} \rightarrow \text{Ag}^+ + \text{e}^-$  oxidation

cell R:  $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$  oxidation

b) Mol cu =  $0.81/64 = 0.0127$  mol

Mol Ag =  $2.7/108 = 0.025$  mol

Since all 3 cells were connected in series, the total  $\text{e}^-$  flow through the circuit are the same for the 10 minutes regardless of the concentration of the

electrolytes. Ag form 1 + ions and Cu form 2+ ions, so the mole ratio is about Ag:Cu=2:1 error due to experimental random error.

c) In cell-R

Initial reaction  $2\text{Al} \rightarrow 2\text{Al}^{3+} + 6\text{e}^-$  at Anode

$6\text{H}^+ + 6\text{e}^- \rightarrow 3\text{H}_2$  at cathode

As the reaction progress, more  $\text{H}^+$  was used up leading to a weak alkaline solution with  $\text{OH}^-$  ions concentrated at anode which combine with  $\text{Al}^{3+}$  to form  $\text{Al}(\text{OH})_3$  ppt which is Amphoteric but insoluble in weak base.  $\text{Al}(\text{OH})_3$  form an insulation layer at the anode preventing further reaction.

A8:

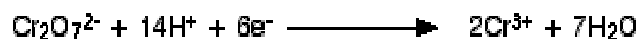
a)  $\text{C}_4\text{H}_8 + \text{H}_2 \rightarrow \text{C}_4\text{H}_{10}$  butane

$\text{C}_4\text{H}_8 + \text{Br}_2 \rightarrow \text{C}_4\text{H}_8\text{Br}_2$  bromination, solution from reddish brown to colorless

$\text{C}_3\text{H}_7\text{OH} + \text{K}_2\text{Cr}_2\text{O}_7$  (act as oxidizing agent)  $\rightarrow \text{C}_3\text{H}_5\text{COOH}$  Propanoic acid

The oxidising agent used in these reactions is normally a solution of sodium or potassium dichromate(VI) acidified with dilute sulphuric acid. If oxidation occurs, the orange solution containing the dichromate(VI) ions is reduced to a green solution containing chromium(III) ions.

The electron-half-equation for this reaction is



## Primary alcohols

Primary alcohols can be oxidised to either aldehydes or carboxylic acids depending on the reaction conditions. In the case of the formation of carboxylic acids, the alcohol is first oxidised to an aldehyde which is then oxidised further to the acid.

### *Partial oxidation to aldehydes*

You get an aldehyde if you use an excess of the alcohol, and distil off the aldehyde as soon as it forms.

The excess of the alcohol means that there isn't enough oxidising agent present to carry out the second stage. Removing the aldehyde as soon as it is formed means that it doesn't hang around waiting to be oxidised anyway!

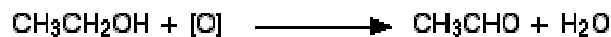
If you used ethanol as a typical primary alcohol, you would produce the aldehyde ethanal, CH<sub>3</sub>CHO.

The full equation for this reaction is fairly complicated, and you need to understand about electron-half-equations in order to work it out.



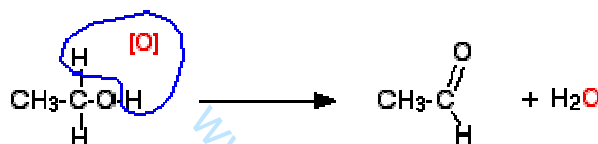
In organic chemistry, simplified versions are often used which concentrate on what is happening to the organic substances. To do that, oxygen from an oxidising agent is represented as [O]. That would produce the much simpler equation:





This means "oxygen from an oxidising agent".

It also helps in remembering what happens. You can draw simple structures to show the relationship between the primary alcohol and the aldehyde formed.

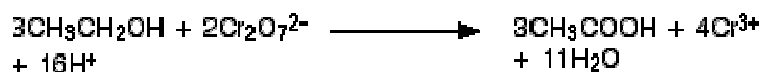


### **Full oxidation to carboxylic acids**

You need to use an excess of the oxidising agent and make sure that the aldehyde formed as the half-way product stays in the mixture.

The alcohol is heated under reflux with an excess of the oxidising agent. When the reaction is complete, the carboxylic acid is distilled off.

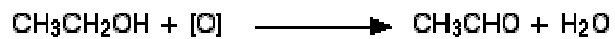
The full equation for the oxidation of ethanol to ethanoic acid is:



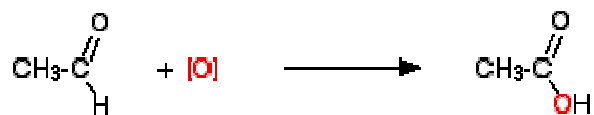
The more usual simplified version looks like this:



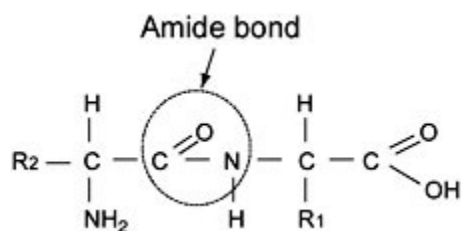
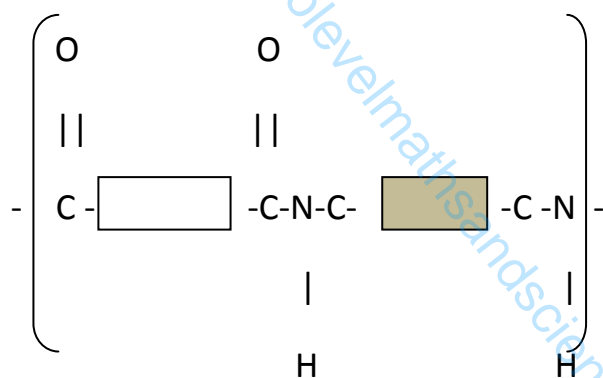
Alternatively, you could write separate equations for the two stages of the reaction - the formation of ethanal and then its subsequent oxidation.



This is what is happening in the second stage:



b) Condensation polymerization, H<sub>2</sub>O will be eliminated from the polymers



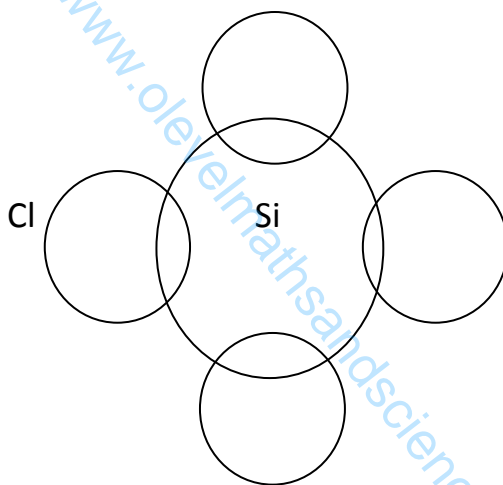
B9:

a) The MP increase from Group 1 to 4 (Si form giant molecular compound), group 5 to 8 are non-metallic forming covalent compound with weak intermolecular bonding thus lower MP.

First 4 element all have MP above room temperature thus are solid at rtp condition.

b) First 2 chloride are ionic compound with strong electrovalent bonding thus high BP. Last 2 chloride are covalent bonded molecules, thus low BP.

SiCl<sub>4</sub> are covalent compound:



B10:

a) Fats (Lipids) & Oil are organic compound composed of C, H, O, N

Unsaturated refer to the presence of C=C bonds which are reactive.

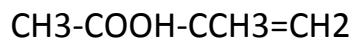
Mole of hydrocarbon =  $10/800 = 0.0125$  mol    mole of H<sub>2</sub> =  $3.6/24 = 0.15$  mol

⇒ Mole of H<sub>2</sub>/mol of hydrocarbon =  $0.15/0.0125 = 12$  mol → 12x C=C bond per molecules.

b) Ester → C<sub>5</sub>H<sub>8</sub>O<sub>2</sub> (No simpler form of empirical formula)

In structural form : CH<sub>2</sub>=CCH<sub>3</sub>-COO-CH<sub>3</sub>

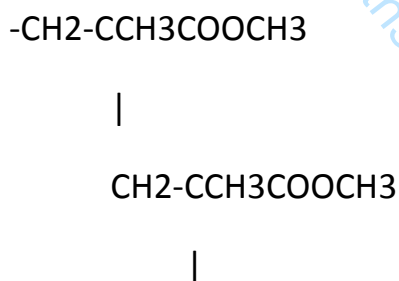
i) Isomer which contain the ester group as:



iii) 2CH<sub>3</sub>CH<sub>2</sub>CH=CHCOOH + Mg → (CH<sub>3</sub>CH<sub>2</sub>CHCHCOOH)<sub>2</sub>Mg + H<sub>2</sub> (g)

iv) CH<sub>2</sub>=CCH<sub>3</sub>COOH + CH<sub>3</sub>OH → CH<sub>2</sub>=CCH<sub>3</sub>-COO-CH<sub>3</sub> + H<sub>2</sub>O by Condensation reaction

v) Addition polymerization by opening of the C=C double bonds

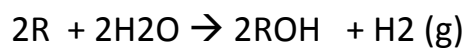


B11: Stability of oxide → Oxides of elements with strong electron affinity will be stable

From the 4 experiments, P less reactive than Cu, Q & R More reactive than Cu thus causing displacement reaction.

a) Q metal dissolved as ions into the solution causing copper metal to be precipitated out of the solution.

b) reactive metal R react directly with water to form H<sub>2</sub> as follows:



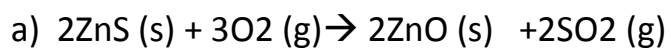
c) In order of ascending reactivity is : P → Cu → Q → R P could be metal like Silver Ag.

d) Metal Q (more reactive and does not easily react with cold water) should be used as the anode.

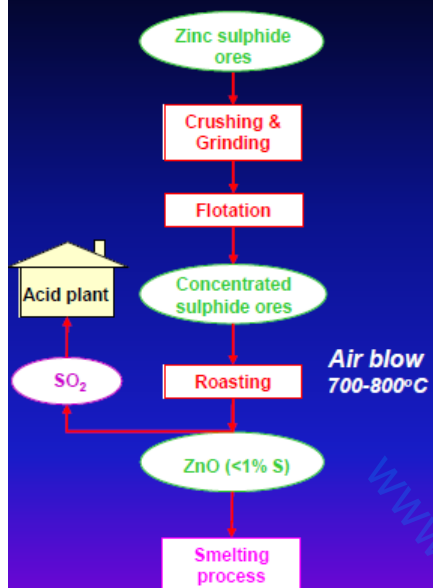
At Anode:  $Q (s) \rightarrow Q^{2+} (aq) + 2e^-$  oxidation

At Cathode:  $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$  reduction

OR B11

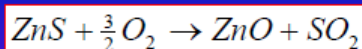


## Pyrometallurgical treatment



### Roasting Process

- The zinc sulphide ores are concentrated by crushing down the size → wet grinding and then flotation.
- **Concentrated zinc sulphide** ores are roasted at  $T \sim 700-800^\circ\text{C}$  with **air blow** to produce **ZnO**.
- The reaction is exothermic, which increases the temp upto  $1000^\circ\text{C}$ .

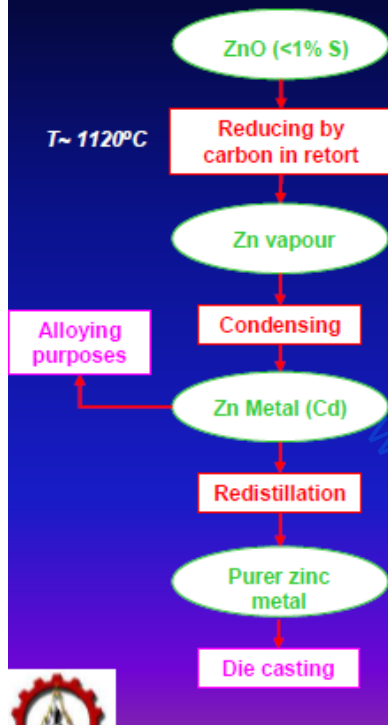


i) Test for presence of SO<sub>2</sub>

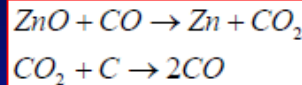
A simple test for sulphur dioxide is to pass the gas over a piece of filter paper soaked in acidified Na<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>. The paper goes from an orange colour for the Cr<sup>6+</sup> to a green colour for Cr<sup>3+</sup>. As the oxidation number of chromium is reduced from 6+ to 3+ this clearly indicates a useful property of sulphur dioxide, namely it is a reducing agent. It has found use as an antioxidant, which can help prevent food from spoiling.

ii)  $2\text{ZnO} + \text{C} \rightarrow 2\text{Zn} + \text{CO}_2$

# Smelting process



• **Zinc oxide** is reduced using **carbon** to obtain **Zn metal**. ( $T \sim 1120^\circ\text{C}$ ),  $T_b \sim 906^\circ\text{C}$



• **Zn vapour** is produced due to high reducing temp and **CO** is released for the use of **preheating** the starting materials.

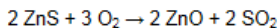
• **Sulphur** must be excluded from the process which can cause the reverse process giving **ZnO** instead.

• **Zn vapour** is then condensed to give a liquid form of **Zn**. (contains small amounts of **Cd**). If used as alloying elements for **Cu** and **Ni** alloys is ok but if used for die casting  $\rightarrow$  redistillation at  $T \sim 765^\circ\text{C}$  to vaporise **Cd** off.

iii) Electrolysis is much more expensive than Carbon reduction method.

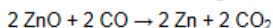
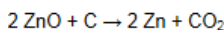
Zinc metal is produced using **extractive metallurgy**.<sup>[79]</sup> After grinding the ore, **froth flotation**, which selectively separates minerals from **gangue** by taking advantage of differences in their **hydrophobicity**, is used to get an ore concentrate.<sup>[79]</sup> A final concentration of zinc of about 50% is reached by this process with the remainder of the concentrate being sulfur (32%), iron (13%), and  $\text{SiO}_2$  (5%).<sup>[79]</sup>

**Roasting** converts the zinc sulfide concentrate produced during processing to zinc oxide.<sup>[78]</sup>

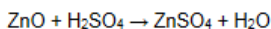


The sulfur dioxide is used for the production of sulfuric acid, which is necessary for the leaching process. If deposits of zinc carbonate, zinc silicate or zinc spinel, like the **Skorpion Deposit** in Namibia are used for zinc production the roasting can be omitted.<sup>[80]</sup>

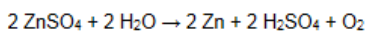
For further processing two basic methods are used: **pyrometallurgy** or **electrowinning**. Pyrometallurgy processing reduces zinc oxide with **carbon** or **carbon monoxide** at  $950^\circ\text{C}$  ( $1,740^\circ\text{F}$ ) into the metal, which is distilled as zinc vapor.<sup>[81]</sup> The zinc vapor is collected in a condenser.<sup>[78]</sup> The below set of equations demonstrate this process:<sup>[78]</sup>



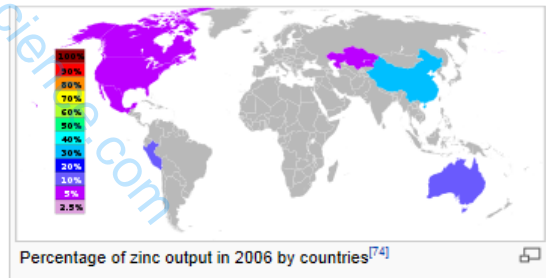
Electrowinning processing leaches zinc from the ore concentrate by **sulfuric acid**.<sup>[82]</sup>



After this step **electrolysis** is used to produce zinc metal.<sup>[78]</sup>



The sulfuric acid regenerated is recycled to the leaching step.



b) Bubbles form at Zn electrode (act as cathode which attracted  $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$  (g)), thus Q must be the Anode which is more reactive than Zn

c) The Zinc block serves as the Anode (sacrificial metal) which dissolves into the seawater (serves as electrolyte) thus protecting the less reactive Steel from corrosion.

d)  $\text{CuSO}_4 (\text{aq}) + \text{Zn} (\text{s}) \rightarrow \text{ZnSO}_4 (\text{aq}) + \text{Cu} (\text{s})$  displacement reaction will occur.

Galvanization is the process of applying a protective zinc coating to steel or iron, in order to prevent rusting. The term is derived from the name of Italian scientist Luigi Galvani.

Although galvanization can be done with electrochemical and electro-deposition processes, the most common method in current use is hot-dip galvanization, in which steel parts are submerged in a bath of molten zinc.



End of paper

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