The Acidic Environment

PART 1

Indicators

All of the substances used in daily life are acidic, basic or neutral, eg. vinegar, lemon juice, aspirin (acid); ammonia, sodium bicarbonate, drain cleaner (basic).
Simple indicators demonstrate the nature of substances over a range of pH by colour change. Please see diagram below.

Indicator →	**Highly acidic**	**Slightly acidic**	**Neutral**	**Slightly alkaline**	**Highly alkaline**
Methyl orange	red	Yellow			
Bromothymol blue		Yellow	blue		
Litmus		Red	blue		
Phenolphthalein				colourless	red

Everyday uses of acids:

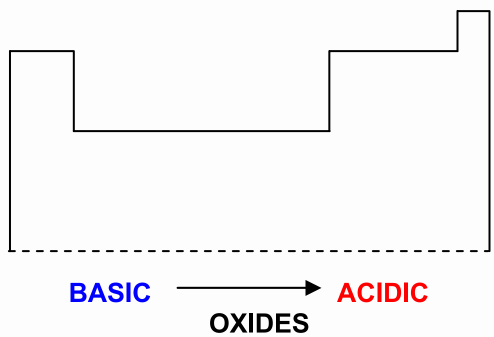
- acetic acid as vinegar for a variety of uses eg. Salad dressing
- swimming pools (approx neutral)
- HCl used to clean bricks

Natural indicators such as the above can be created by taking a relevant plant, crushing it into a powder if needed, and dissolving in water/methanol, etc to extract the dye. (To make indicator paper, soak a piece of filter paper in desired indicator.)

Oxides of carbon, nitrogen and sulfur.

An **acidic oxide** react with water to make an acid, or will react with bases to make salts (some oxides of non-metals).

Basic oxides react with acids to form salts, but not react with alkali solutions).
The acidity of the oxide of an element increases from left to right on the periodic table. Basic oxides become more basic moving down a group, acidic oxides become more acidic up a group. Oxides of Aluminium and Beryllium are **amphoteric** (do not confuse this term with amphiporitc!!)



Chemical equilibrium occurs when a reaction does not go to completion, rather the products cause the back reaction to occur at the rate of the forward reaction. The relative amount of product(s) and reactant(s) at equilibrium can be affected by altering the equilibrium in order to favour one side.

- **Le Chatelier's principle** states that if equilibrium is disturbed, the system reacts to **minimise this disturbance**.

Factors that can affect the equilibrium include:

- Concentrations of reactant/product involved in the reaction (when concentration of a species is altered the equilibrium reacts by decreasing the concentration, shifting the equilibrium to the side that removes this species)
- Pressure in a reaction involving gases (eg if pressure is increased the equilibrium will shift to the side that makes the least number of moles)
- Temperature (it is essential to know if the reaction is exo/endothermic)

Solubility of CO_2 in water:

$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$ $\Delta H<0$

Disturbance	Effect on equilibrium (right means inc. solubility)
Increased pressure	Favours reaction with the least moles of gas in product. This reduces pressure, shifts equilibrium right, more CO_2 dissolves.
Increase in concentration of CO_2 in system	System reduces concentration of CO_2 and thus it alters, ie, shifts equilibrium right
Increase in temperature	Favours endothermic reaction (that which will absorb the increase in temp), ie reverse reaction, ie shifts left, and CO_2 is **less** soluble

Natural and industrial sources of SO_2 and nitrogen oxides

- **Sulfur dioxide (SO_2)**
  - Geothermal hot springs, volcanic hot spots
  - Processing or combustion of fossil fuels ($S(compounds) + O_2(g) \rightarrow SO_2(g)$)
  - Extraction of metals from sulfide ores
    (eg $2ZnS(g) + 3O_2(g) \rightarrow 3ZnO(s) + 2SO_2(g)$)

- **Oxides of nitrogen**: nitrous oxide (N_2O), nitric oxide (NO), nitrogen dioxide (NO_2)
  - Lightning in the atmosphere ($O_2(g) + N_2(g) \rightarrow 2NO(g)$)
  - Combustion chambers (cars, power stations) same result
  - A further reaction of $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
  - N_2O formed by bacteria (from nitrogenous fertiliser, contributes to greenhouse effect)
  - Mixture of NO and NO_2 referred to as NO_x

Pollutant oxides **can become dissolved** in water droplets, within large volumes of air in the atmosphere.

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$ ……………………………………………….(1)

$2H_2SO_3(aq) + O_2(g) \rightarrow 2H_2SO_4(aq)$ (catalysed) ……..……………………..(2)

$2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$ ………………………………(3)

$2HNO_2(aq) + O_2(g) \rightarrow 2HNO_3(aq)$ (catalysed) …………………………(4)

- at $0^{o}C$ and 100kPa pressure, molar volume of gases is 22.71 L
- at $25^{o}C$ and 100kPa pressure, molar volume of gas is 24.79 L

High SO_2 and NO_x emission cause acidic rain (eqn 2, 3, 4): **acid rain** is rain with a higher H^+ concentration than normal, > 10^{-5} M. Ordinary rain unaffected by these pollutants contains some acidic carbonic acid (from CO_2) and usually contains $10^{-6} – 10^{-5}$ M $[H^+]$.

Acid rain causes:

- increasing lake acidity (affects fish populations)
- damage to pine forests and vegetation
- erosion of marble and limestone buildings/statues (these contain carbonates which react with acids)

Properties of SO_2 and NO_x

SO_2	N_2O	NO	NO_2
Colourless Pungent odour Soluble	Colourless Sweet smell Insoluble	Colourless No smell Insoluble	Reddish-brown Choking odour Soluble in water
• food preservative • bleaching • fumigant	• anaesthetic • propellant	• synthesizing nitric oxide	• nitric acid • fertilisers • explosives
• Breathing difficulties at 1ppm			• Breathing difficulties at 5ppm • Leads to forming O_3

Acids in food, drinks and within our stomachs

Arrhenius defined acids as able to form H^+ in water (ionisation reaction). The H^+ joins with water to make the hydronium ion (H_3O^+).
According to the Lowry-Brönsted definition (1923), acids donate protons and bases are accept protons.

Naturally occurring acids	Manufactured acids
– Acetic acid (ethanoic acid) – vinegar – Citric acid (2-hydroxypropane-1,2,3-tricarboxylic acid, $C_6H_8O_7$) – Vitamin C (ascorbic acid, $C_6H_8O_6$) – Hydrochloric acid (in our stomachs)	– sulfuric acid – hydrobromic acid – nitric acid – phosphoric acid

The pH scale is a valuable scale for comparing the concentrations of acids and alkalis, since the concentrations of the ions themselves involved range over a very large linear scale. It is more convenient to use a smaller, everyday language scale, and thus pH is used.

pH = -log [H_3O^+] A change by 1 in pH is a tenfold increase in [H^+].

Note: the number of decimal places for pH should equal the number of significant figures for [H_3O^+]

TIPS FROM AN EXPERIENCED MARKER

- students must apply knowledge, verbs indicate depth of knowledge
- know your verbs – analyse, discuss, evaluate, assess, compare, describe, explain, justify
- Large volumes of work will not ensure good marks, if you contradict yourself you will lose marks don't write 2 points unless you are sure they are both correct.
- For questions over 3 marks it is a good idea to write a plan
- show states in equations
- Calculations must make sense, must show some logical sequence, answers must pay heed to significant figures and units. Don't round off at each step. Round off your answer to the correct number of significant figures. You must show working as marks are allocated for each step.
- Always make reference to the stimulus material that may be included. If the question says use the table, THEN, use the table!!!!
- Check you calculated answer makes sense eg a pH of 995 does not make sense or a base with a pH of 3.5?
- use dot points to answer questions and not long winded sentences
- care with plurals if there is an S on the end of the word examples it means more than one example
- USE EQUATIONS THIS IS WHAT WE DO IN CHEMISTRY AND MAKE SURE THEY ARE BALANCED
- If appropriate draw a diagram to support your answer. For questions about practicals, immediately draw a diagram and make it a decent labelled effort done in pencil.
- If asked about a practical you should include the equipment used, the relative amounts of reagents and any controls you implemented. Don't just say what you did – you must say how you did it.
- Plan you time and approach to the paper. If you finish more than ½ to ¾ early you haven't made a decent effort, so go back and see where you can improve on your answers

GOOD LUCK!!!