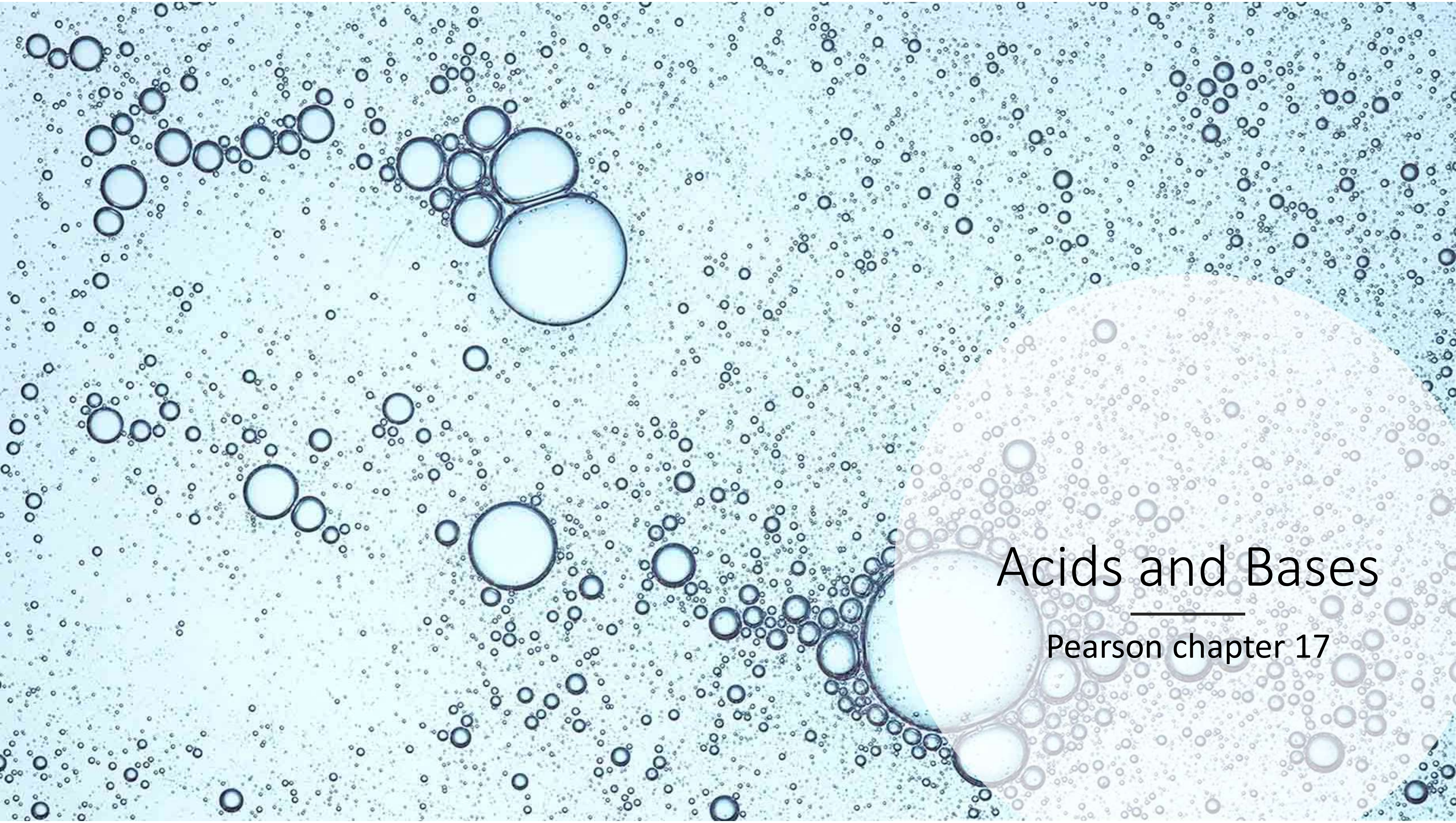
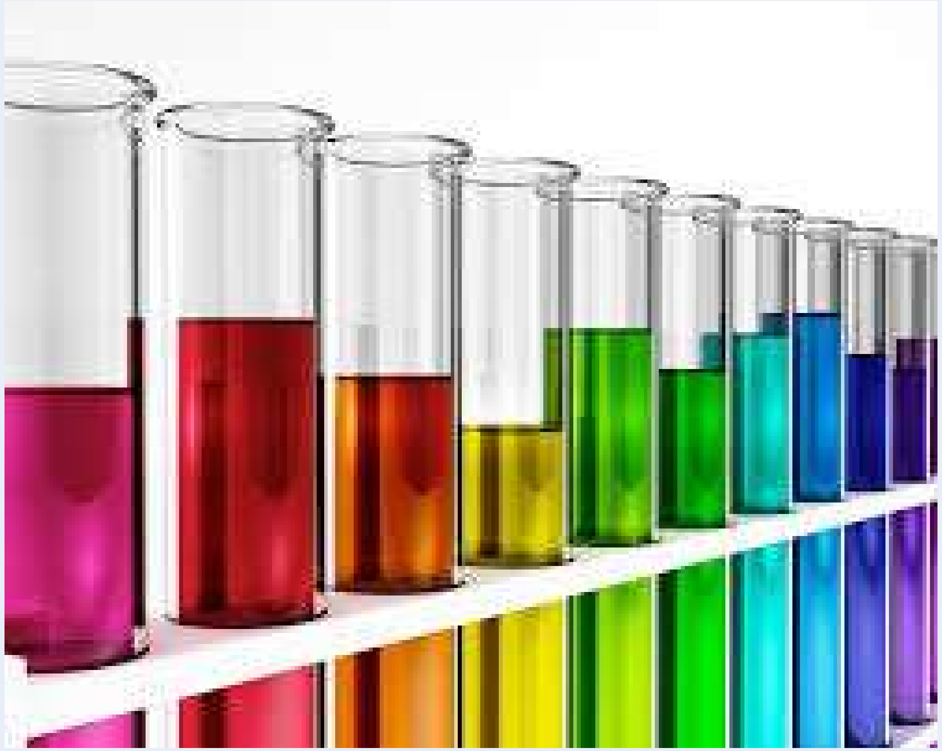


Acids and Bases

Pearson chapter 17

Outcomes

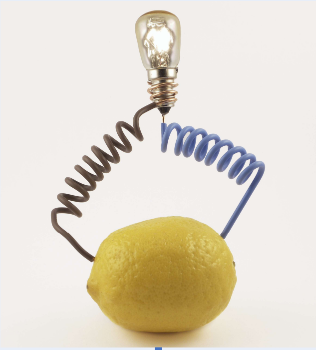
- Properties of acids and bases
- Indicators
- The Arrhenius theory of Acids and Bases
- Strong and weak acids and bases

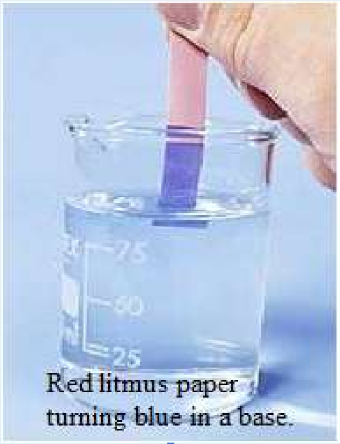


Common acids and bases

	Acid		Base	
common examples of **strong** forms	HCl	hydrochloric acid	$LiOH$	lithium hydroxide
	HNO_3	nitric acid	$NaOH$	sodium hydroxide
	H_2SO_4	sulfuric acid	KOH	potassium hydroxide
			$Ba(OH)_2$	barium hydroxide
some examples of **weak** forms	CH_3COOH and other organic acids	ethanoic acid	NH_3	ammonia
	H_2CO_3	carbonic acid	$C_2H_5NH_2$ and other amines	ethylamine
	H_3PO_4	phosphoric acid		

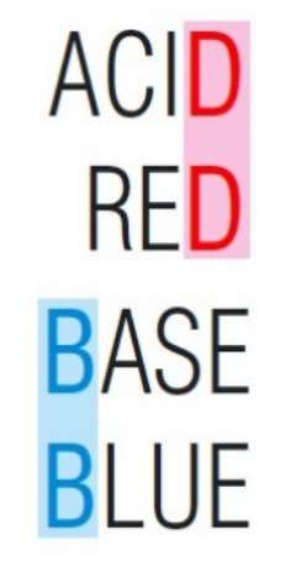
Properties of acids and bases

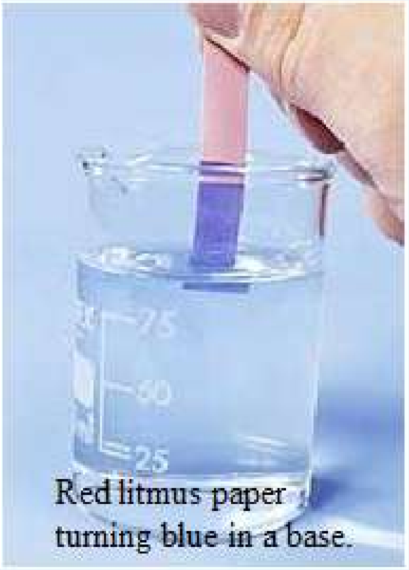




ACIDS	BASES
Conduct electricity	Conduct electricity
Turn blue litmus paper red	Turn red litmus paper blue
Sour taste	Bitter taste
Corrosive (burn)	Slippery, soapy feel

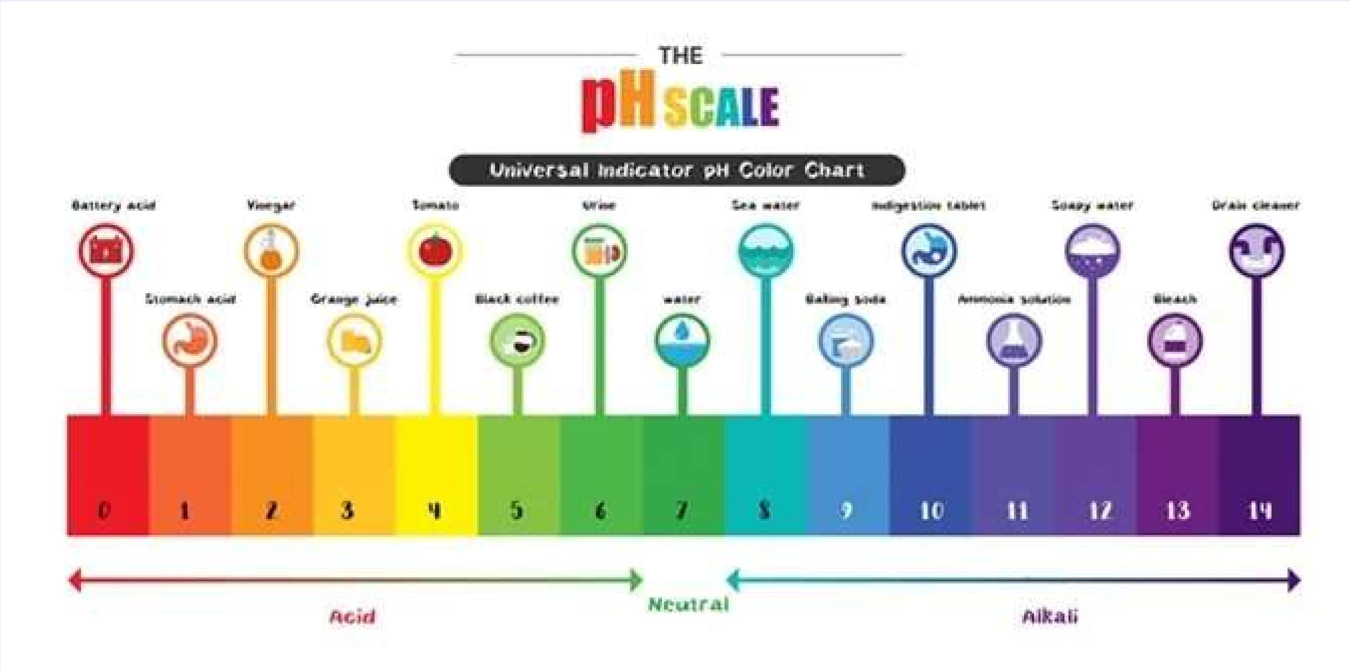
Properties of acids and bases





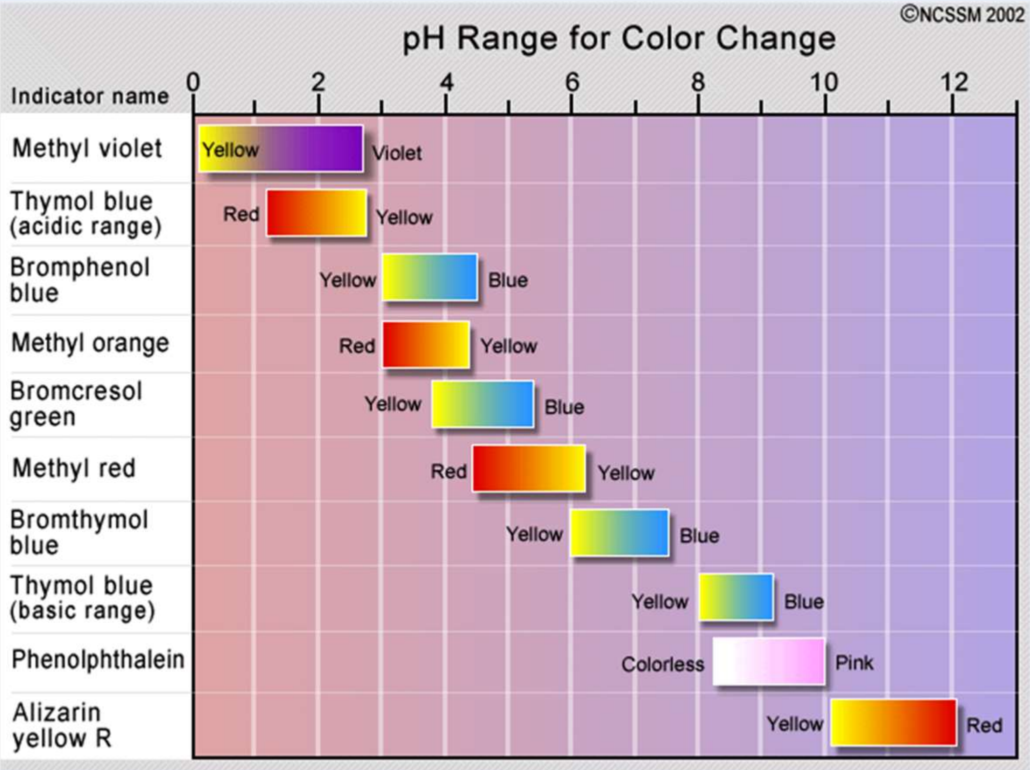
The colour of hydrangea flowers is dependent upon the pH of the soil

pH scale

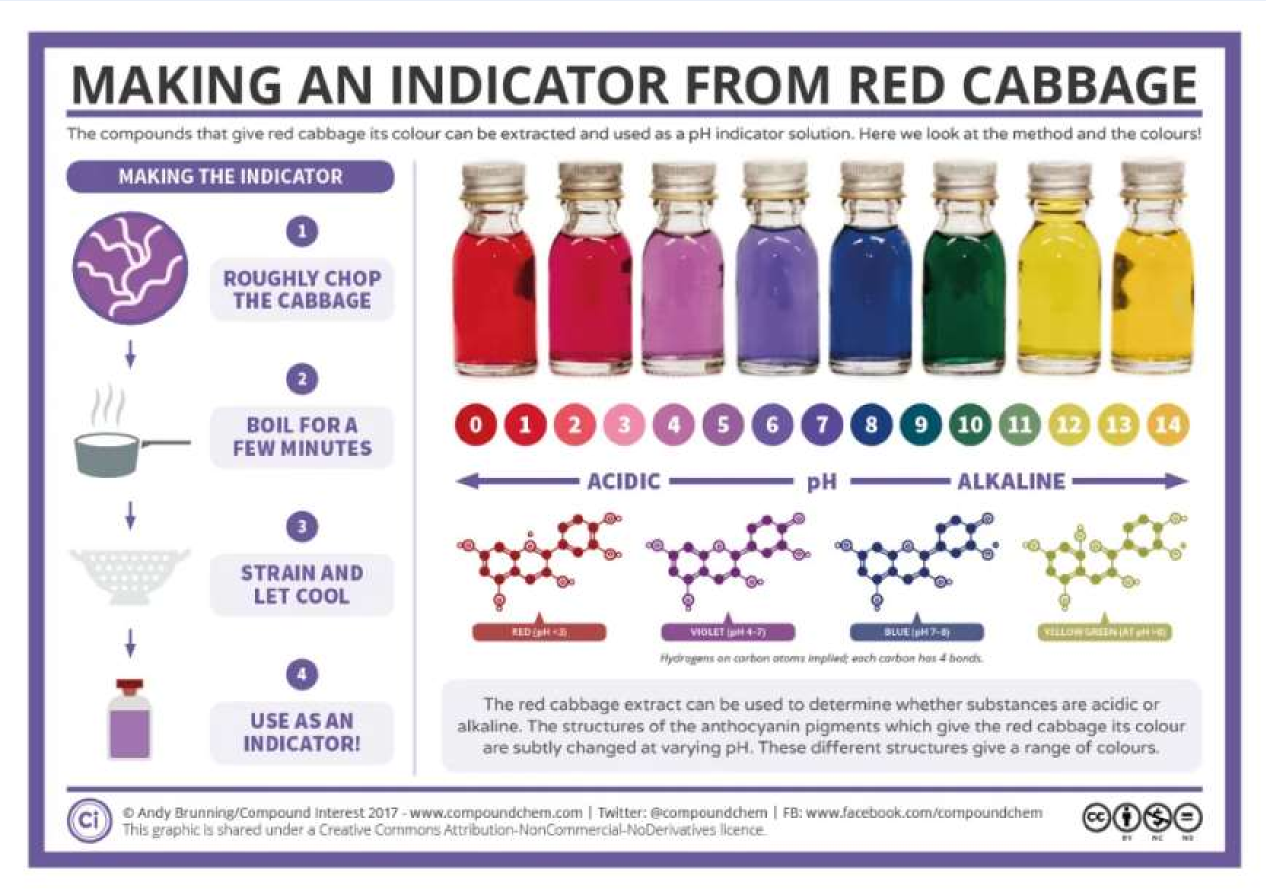

THE
pH SCALE
Universal Indicator pH Color Chart
Battery acid
Stomach acid
Vinegar
Orange juice
Tomato
Black coffee
water
Sea water
Baking soda
Ammonia solution
Soapy water
Bleach
Drain cleaner
0
1
2
3
4
5
6
7
8
9
10
11
12
13
14
Acid
Neutral
Alkali

Indicators

- Indicators are coloured compounds which change colour based on the pH of the solution



Indicators



MAKING AN INDICATOR FROM RED CABBAGE
The compounds that give red cabbage its colour can be extracted and used as a pH indicator solution. Here we look at the method and the colours!
MAKING THE INDICATOR
1 ROUGHLY CHOP THE CABBAGE
2 BOIL FOR A FEW MINUTES
3 STRAIN AND LET COOL
4 USE AS AN INDICATOR!
0 1 2 3 4 5 6 7 8 9 10 11 12 13 14
ACIDIC pH ALKALINE
RED (pH <3)
VIOLET (pH 4-7)
BLUE (pH 7-8)
YELLOW GREEN (AT pH >8)
Hydrogens on carbon atoms implied; each carbon has 4 bonds.
The red cabbage extract can be used to determine whether substances are acidic or alkaline. The structures of the anthocyanin pigments which give the red cabbage its colour are subtly changed at varying pH. These different structures give a range of colours.
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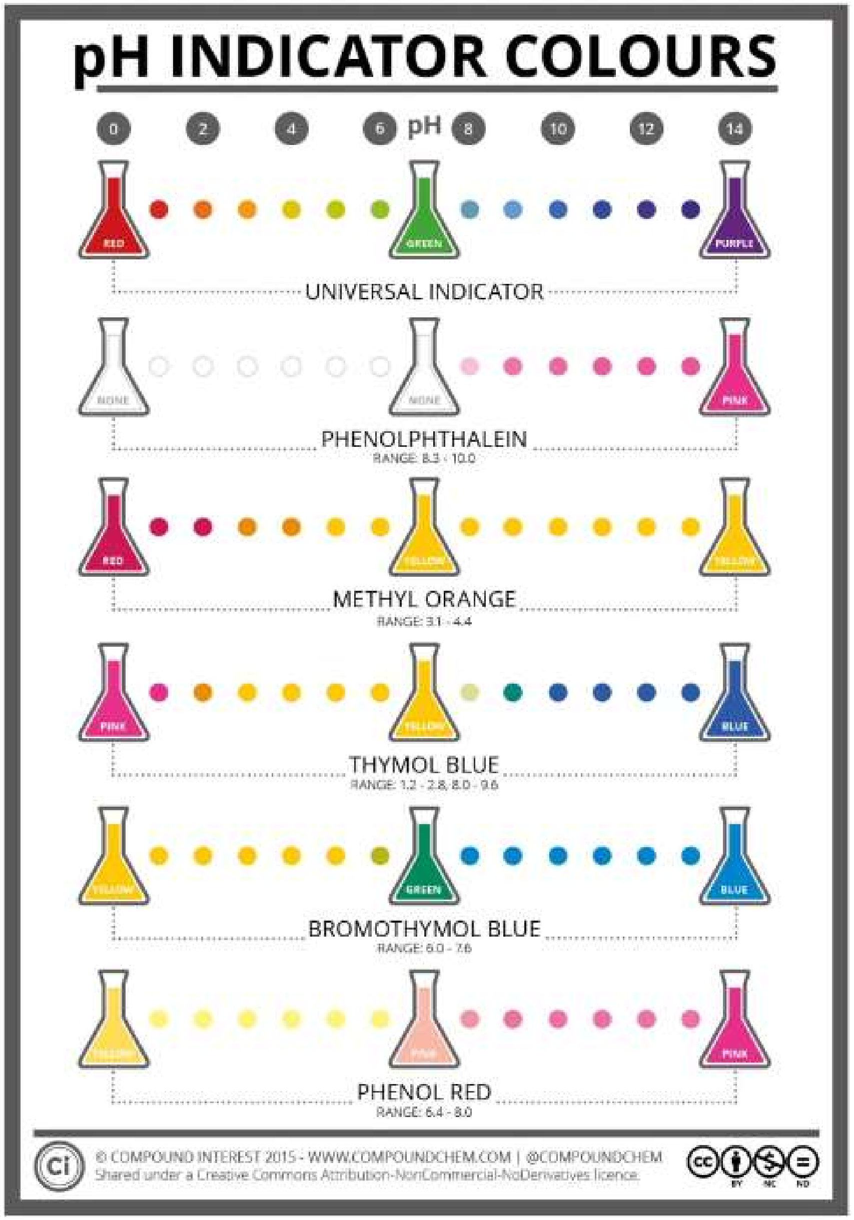

pH INDICATOR COLOURS
0 2 4 6 pH 8 10 12 14
RED GREEN PURPLE
UNIVERSAL INDICATOR
NONE NONE PINK
PHENOLPHTHALEIN
RANGE: 8.3 - 10.0
RED YELLOW YELLOW
METHYL ORANGE
RANGE: 3.1 - 4.4
PINK YELLOW BLUE
THYMOL BLUE
RANGE: 1.2 - 2.8, 8.0 - 9.6
YELLOW GREEN BLUE
BROMOTHYMOL BLUE
RANGE: 6.0 - 7.6
YELLOW PINK PINK
PHENOL RED
RANGE: 6.4 - 8.0
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Ionisation of water

- Water (H_2O) is covalent molecular
- As we learnt previously, some covalent molecular compounds (such as ethanoic acid) ionise to form ions in aqueous solutions – we call these weak electrolytes.
- Water, to a small extent, self-ionises into $H^+_{(aq)}$ and $OH^-_{(aq)}$

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)} \qquad 2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

Recall: The double arrow means the process is reversible

Ionisation of water

$$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

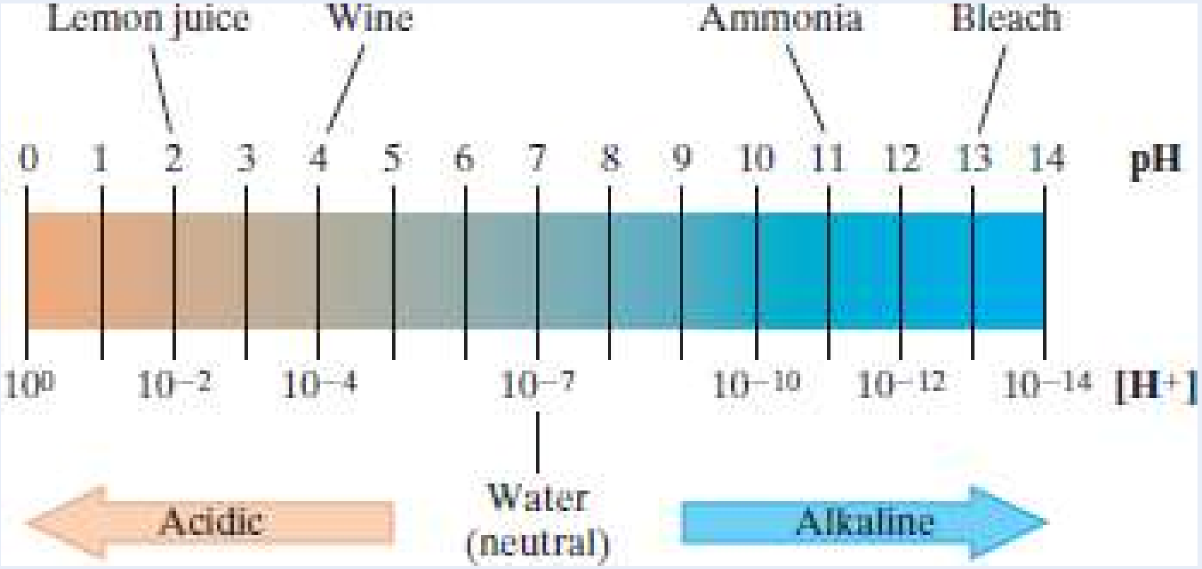
- In pure water, we will form an equal amount of $H^+_{(aq)}$ and $OH^-_{(aq)}$
- For this reason we say pure water is neutral.
- The ions rapidly recombine so the concentrations of the ions is low, the concentration of $H^+_{(aq)}$ and $OH^-_{(aq)}$ in pure water is 1.0×10^{-7} mol L^{-1}
- Concentration of $H^+_{(aq)}$ can be used to calculate pH of a solution

$$\mathrm{pH} = -\log[\mathrm{H}^+]$$

If	$[H^+] = 1.0 \times 10^{-7}$ mol L^{-1}
Then	pH = 7

pH scale

The pH scale is a negative logarithmic scale. The logarithmic part means that pH changes by 1 unit for every factor of 10 change in concentration of H^+. The negative sign in front of the log tells us that there is an *inverse relationship* between pH and $[\mathrm{H}^+]$: when pH increases, $[\mathrm{H}^+]$ decreases, and vice versa.



Arrhenius theory of acids and bases

- An acid has H in its formula and dissolves in water to form hydrogen ions (H^+)
- A base has OH in its formula and when added to water forms hydroxide ions (OH^-)
- These ions are produced by a reversible reaction known as ionisation or dissociation

Examples

- An acid ionises to produce hydrogen ion

$$HNO_{3(l)} \rightleftharpoons H^+_{(aq)} + NO3^-_{(aq)}$$

- A base dissociated to produce hydroxide ion

$$NaOH_{(s)} \rightleftharpoons Na^+_{(aq)} + OH^-_{(aq)}$$

Arrhenius theory of acids and bases

- Arrhenius theory explains:
- Because all acids have H+ ions they all react with carbonates, hydrogen carbonates, and metals such as magnesium
- Solutions of acids can conduct electricity because of the H+ ions and negative ions
- When mixed together H^+ and OH^- form water H_2O

Arrhenius theory of acids and bases

- Limitations
- Some substances that behave as bases, (produce OH- ions in solution) do not contain OH (NH_3 and carbonates)
- Reactions such as $NH_3 + HCl \rightarrow NH_4Cl$ cannot be explained
- Not all salts are neutral
- H+ ions cannot exist in water for a long time (H_3O^+)

Strong and weak acids and bases

- According to Arrhenius theory, when strong acids are dissolved in water, all the acid molecules break up into hydrogen ions and negative ions
- 1 mole HCl → 1 mole H^+ + 1 mole Cl^-
- When a weak acid is added to water only some of the molecules are ionised
- 1 mole HF → 0.03 mole H^+ + 0.03 mole F^-
- (H^+ are actually H_3O^+)

Strong and weak acids and bases

Strong Acids		Weak Acids	
Nitric acid	HNO_3	Acetic acid	CH_3COOH
Sulfuric acid	H_2SO_4	Hydrofluoric acid	HF
Hydrochloric acid	HCl	Phosphoric acid	H_3PO_4
Hydrobromic acid	HBr	Sulfurous acid	H_2SO_3
Hydroiodic acid	HI	Ammonium ion	NH_4^+
Perchloric acid	$HClO_4$	Hydrogen sulfate ion	HSO_4^-
Strong Bases		**Weak Bases**	
Metal hydroxides	$NaOH$	Ammonia	NH_3
Metal oxides	Na_2O	Hydrogen carbonate	
		Carbonate ion	
		Phosphate ion	

Polyprotic acids

- Monoprotic acid – each molecule of acid produces one hydrogen ion only during ionisation

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

- Polyprotic acids:
  - produce more than one hydrogen ion when dissolved in water

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

  - Hydrogen sulfate produced acts as a weak acid in the next process

$$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$$

On going work

- Pearson chapter 17.1 and 17.2
- STAWA Sets
  - pH scale calculations SET 34
  - Acid and Base reaction stoichiometry SET 36