Unit 1
Inorganic and Physical Chemistry
Chemical Equilibrium
Dynamic Equilibrium

Revision

Dynamic equilibrium happens in a closed system when the rate of reverse reaction equals rate of forward reaction. The composition of reactants and products appears to stay constant.

Homogenous: all species in the same phase

Heterogenous: species in more than one phase

Equilibrium constant

The equilibrium constant ($K$) is expressed in terms of concentration (or pressure for gases).

For the reaction:

$$aA + bB \rightarrow cC + dD$$

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Concentrations of pure solids or liquids are constant and are given the value 1, thus

$$S(s) + O_2(g) \rightarrow SO_2(g) \quad K = \frac{[SO_2]}{[O_2]}$$

Equilibrium constant is for a given temperature and is independent of concentration, pressure or the presence of a
catalyst. The effect of changes in concentration can be explained in terms of a fixed equilibrium constant. $K$ has no units - it is dimensionless.

e.g. When dilute hydrochloric acid is added to a yellow solution of sodium chromate the solution turns orange:

$$2\text{CrO}_4^{2-}(aq) + 2\text{H}^+(aq) \rightarrow \text{Cr}_2\text{O}_7^{2-}(aq) + \text{H}_2\text{O}(l)$$

Yellow  orange

The added $\text{H}^+(aq)$ ions increase the rate of the forward reaction compared to the reverse reaction. This continues until the rates of the forward and reverse reactions are the same and equilibrium is re established.

The position of the equilibrium has moved to the right. Overall the increase in the concentration of $\text{H}^+(aq)$ is compensated for by a decrease in the concentration of $\text{CrO}_4^{2-}$ ions and an increase in the concentration of $\text{Cr}_2\text{O}_7^{2-}$ ions and $\text{H}_2\text{O}$ molecules, thus keeping the value of $K$ the same:

$$K = \frac{[\text{Cr}_2\text{O}_7^{2-}]}{[\text{CrO}_4^{2-}]^2[\text{H}^+]^2}$$

This fact is stated formally in Le Chatelier's principle, which states that 'when a reaction at equilibrium is subjected to change, the composition alters in such a way as to minimise the effects of the change'.
The effect of changing temperature

The effect of a temperature change on a reaction in equilibrium can be demonstrated in the laboratory using the following example:

\[ \text{N}_2\text{O}_4 (g) \rightleftharpoons 2\text{NO}_2 (g) \quad \Delta H = +ve \]

yellow \quad brown

Three boiling tubes filled to an equal colour intensity with this gas mixture are collected at room temperature. One is placed in ice water (~0°C) and one in hot water (~80°C) while the third is kept, as a control, at room temperature (Figure 3). The three tubes are left for about five minutes and the colour intensities are then compared.

From the colour changes that have taken place, it is evident that the relative concentrations of N\(_2\)O\(_4\) and NO\(_2\) have been changed by a change in temperature. This means that the actual value of \(K\) has changed. Equilibria are therefore temperature dependent. In fact most equilibrium constants are quoted at a specific temperature.

The equilibrium constant is:

\[ K = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} \]
The reaction is endothermic for the forward reaction and exothermic for the reverse. The colour change from 20°C to 0°C shows that more N₂O₄ is formed at 0°C and therefore [N₂O₄] increases and [NO₂] decreases. This leads to a fall in the value of $K$.

Similarly, the colour change from 20°C to 80°C shows that more NO₂ is formed at 80°C and therefore [NO₂] increases and [N₂O₄] decreases. This leads to a rise in the value of $K$. These observations can be related to the $\Delta H$ value and are summarised in the following way.

*For endothermic reactions a rise in temperature causes an increase in $K$ while for exothermic reactions a rise in temperature causes a decrease in $K$.***

**K and Equilibrium Position**

A high value of $K$ means that the fraction used to calculate it is 'top heavy' so there is more product, i.e. equilibrium lies to the right. The table below summarises this for different reactions:

<table>
<thead>
<tr>
<th>System</th>
<th>Value of $K$</th>
<th>Position of equilibrium</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag⁺(aq) + 2NH₃(aq) $\rightleftharpoons$ [Ag(NH₃)₂]⁺(aq)</td>
<td>$1.7 \times 10^7$ at 25°C</td>
<td>Because $K&gt;&gt;1$ the equilibrium lies to the right</td>
</tr>
<tr>
<td>CH₃COOH(aq) $\rightleftharpoons$ CH₃COO⁻(aq) + H⁺(aq)</td>
<td>$1.8 \times 10^{-5}$ at 25°C</td>
<td>Because $K&lt;&lt;1$ the equilibrium lies to the left</td>
</tr>
<tr>
<td>N₂O₄(g) $\rightleftharpoons$ 2NO₂(g)</td>
<td>0.87 at 55°C</td>
<td>Because $K=1$ the equilibrium lies to neither the left nor the right</td>
</tr>
</tbody>
</table>
Effect of a Catalyst

Remember that all a catalyst does is lower activation energy. This will be true for forward and reverse reactions, so a catalyst speeds up the rate of establishment of equilibrium. It doesn’t affect the position of equilibrium; therefore it must have no effect on K.

Questions:

1. The value for the equilibrium constant, K, for a specific example of the following reaction is equal to 1:

   \[ \text{acid + alcohol } \rightleftharpoons \text{ester + water} \]

   a) i) Predict the maximum yield of ester, given this value of K.

   ii) Give one reason why this yield might not be achieved in practice.

   b) A student suggested that a careful choice of catalyst could increase the yield of ester. Comment on this suggestion.

   c) Another student suggested that increasing the concentration of the alcohol in the reaction mixture would increase the yield of ester by altering the value of K. Comment on this suggestion.

2. The balanced equation for the principle reaction in the contact process is:

   \[ 2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g) \quad \Delta H = -\text{ve} \]

   a) Write an expression for K, the equilibrium constant for this reaction, when the system has reached equilibrium.

   b) i) State Le Chatelier’s Principle.
ii) Use this generalisation to predict how an increase in temperature would affect the position of equilibrium.

iii) Explain the effect this increase in temperature would have on the value of K.
**Acid/Base Equilibria**

From the National 5 course it will be remembered that the pH of an aqueous solution is a measure of the concentration of hydrogen ions in the solution. In fact the pH of any aqueous solution can be calculated using the expression:

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

The following examples illustrate its use:

- Pure water \([H^+] = 10^{-7}\) so pH = 7
- 1.0 mol l\(^{-1}\) HCl \([H^+] = 1.0 \times 10^0\) so pH = 0
- 0.2 mol l\(^{-1}\) HCl \([H^+] = 0.2 \times 10^{-1}\)

So pH = \(-\log 2 - \log (10^{-1})\)

\[ = -0.3 + 1.0 = 0.7 \]

For calculations involving strong acids and alkalis it can be assumed that they are 100% dissociated and that the small number of hydrogen ions supplied by the water can be ignored.

In S4, acids were defined as producing hydrogen ions in solution and bases as producing hydroxide ions in solution. These two ions combine in the neutralisation equation:

\[ H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \]

However, this definition is simplistic and can’t explain all acids. For example, pure hydrogen chloride is a gas that contains no \(H^+\) ions and ammonia has no \(OH^-\) ions but can neutralise an acid. It was also discovered that the \(H^+\) ion could not exist in aqueous solution. The \(H^+\) ion is so small that the electric field it
creates is huge and in water it attracts a lone pair of electrons to form $\text{H}_3\text{O}^+$, which is called the hydronium ion.

$$\text{H}^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq})$$

The shorthand $\text{H}^+(\text{aq})$ is always used in stoichiometric and equilibrium equations, although this is not strictly accurate. In 1923 the new discoveries led two chemists, Brønsted and Lowry, to define acids and bases in a different way:

*An acid is any substance capable of donating a proton.*

*A base is any substance capable of accepting a proton.*

**The dissociation of water**

In Brønsted-Lowry terms the ionisation of water is represented by:

$$\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$$

acid + base $\rightarrow$ conjugate acid + conjugate base

The equilibrium constant is:

$$K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]^2}$$

Since water is a liquid, $[\text{H}_2\text{O}] = 1$ (see page 37), giving:
\[ K = [H_3O^+][OH^-] \]

This particular equilibrium constant is known as the **ionic product** \((K_w)\) for water and has the value of \(1.0 \times 10^{-14}\) at 25°C.

The water equilibrium is more usually written as:

\[ H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \quad \text{and} \quad K_w = [H^+][OH^-] \]

In **pure** water, for every molecule which ionises, one \(H^+\) and one \(OH^-\) ion are produced, hence the \([H^+]\) in mol l\(^{-1}\) must equal the \([OH^-]\) in mol l\(^{-1}\), i.e. the number of \(H^+\) and \(OH^-\) ions in water are equal.

Substitution of \([OH^-]\) by \([H^+]\) in the above equilibrium expression gives:

\[ [H^+]^2 = 10^{-14} \text{ mol}^2 \text{ l}^{-2} \]

taking square roots \([H^+] = 10^{-7} \text{ mol l}^{-1}\)

so also \([OH^-] = 10^{-7} \text{ mol l}^{-1}\)

\[ K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \]

This relationship can be used to calculate the pH of an alkali as shown:

0.01 mol l\(^{-1}\) NaOH \([OH^-] = 10^{-2}, [H^+] = 10^{-12}, \text{pH} = 12\)

0.5 mol l\(^{-1}\) NaOH \([H^+] = 2 \times 10^{-14} \text{ pH} = -\log 2 - \log(10^{-14})

= \(-0.3\) + 14 = 13.7
Kw is always quoted as $1.0 \times 10^{-14}$ at 25°C since the value varies with temperature:

$$\text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq) \quad \Delta H^{+ve}$$

As the reaction is endothermic, an increase in temperature moves the equilibrium to the right and a decrease moves it to the left.

Questions

Calculate the pH of the following solutions:
(a) 0.35 mol l$^{-1}$ HNO$_3$
(b) 0.14 mol l$^{-1}$ H$_2$SO$_4$ (assume fully ionised)
(c) 0.78 mol l$^{-1}$ NaOH

**Conjugate Acids and Bases**

When an acid donates a proton the species left is called the conjugate base of that acid. When a base accepts a proton the species formed is called the conjugate acid of that base.

<table>
<thead>
<tr>
<th>acid</th>
<th>+ base</th>
<th>$\rightleftharpoons$</th>
<th>conjugate base</th>
<th>+</th>
<th>conjugate acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>+ H$_2$O</td>
<td>$\rightleftharpoons$</td>
<td>Cl$^-$</td>
<td>+</td>
<td>H$_3$O$^+$</td>
</tr>
<tr>
<td>CH$_3$COOH</td>
<td>+ H$_2$O</td>
<td>$\rightleftharpoons$</td>
<td>CH$_3$COO$^-$</td>
<td>+</td>
<td>H$_3$O$^+$</td>
</tr>
<tr>
<td>H$_2$O</td>
<td>+ NH$_3$</td>
<td>$\rightleftharpoons$</td>
<td>OH$^-$</td>
<td>+</td>
<td>NH$_4^+$</td>
</tr>
<tr>
<td>H$_2$O</td>
<td>+ CH$_3$COO$^-$</td>
<td>$\rightleftharpoons$</td>
<td>OH$^-$</td>
<td>+</td>
<td>CH$_3$COOH</td>
</tr>
</tbody>
</table>

This table also illustrates the **amphoteric** nature of water since it can act both as a proton acceptor and a proton donor.
What is a 'Strong' Acid?

The strongest common acids are HCl, HBr, HI, HNO₃, HClO₃, HClO₄, and H₂SO₄. All strong acids dissociate completely in solution:

\[
\text{HNO}_3(aq) \rightleftharpoons \text{H}^+(aq) + \text{NO}_3^-(aq)
\]

Therefore, the pH depends only on the initial molarity of the acid.

The strongest common bases are oxides and hydroxides of Group I metals. All strong bases dissociate completely in solution:

\[
\text{O}_2^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons 2\text{OH}^-(aq)
\]

Bases do not need to include OH⁻ ions:

\[
\text{H}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{(g)} + \text{OH}^-(aq)
\]

\[
\text{N}^3-(aq) + 3\text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + 3\text{OH}^-(aq)
\]

Therefore, the [OH⁻], and hence the [H⁺] and pH, depends only on the initial molarity of the base.

Dissociation and strength

The dissociation of any acid, HA, in aqueous solution can be represented by the equation:

\[
\text{HA(aq) + H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(aq) + \text{A}^-(aq)
\]

acid base conjugate conjugate acid base
The dissociation constant of acid HA is a measure of the strength of the acid. It is simply the equilibrium constant from the above equation and is:

\[ K_a = \frac{[H_3O^+][A^-]}{[HA]} \]

The [H₂O] has been omitted as it is the solvent and taken to be 1.

In **strong acids** the above equilibrium lies to the right with effectively complete dissociation so that \( K_a \) has no meaning. However, some acids do not dissociate well (less than 5%) in aqueous solution and these are called **weak acids**.

Conveniently, nearly all acids fall into one of these two categories. Within the weak acids, the value of \( K_a \) gives a measure of how weak (or how dissociated) the acid is: the smaller the value of \( K_a \), the weaker the acid. A similar constant, \( K_b \) exists for the base equilibrium, but in general \( K_a \) is used to calculate pH for both acids and bases - we will see how this is done later.
pH of Salt Solutions

The pH of a salt solution depends on the strength of the acid and base from which it was formed.

<table>
<thead>
<tr>
<th>Acid</th>
<th>Base</th>
<th>pH of salt</th>
</tr>
</thead>
<tbody>
<tr>
<td>Strong</td>
<td>Weak</td>
<td>Acidic</td>
</tr>
<tr>
<td>Strong</td>
<td>Strong</td>
<td>Neutral</td>
</tr>
<tr>
<td>Weak</td>
<td>Strong</td>
<td>Alkaline</td>
</tr>
</tbody>
</table>

All salts are strong electrolytes and completely dissociate in solution. The ions of a weak acid or base when dissolved will set up an equilibrium with the ions in water.

**Acidic Salt**

In water you have:

\[
\text{H}_2\text{O} \quad \text{(l)} \quad \rightleftharpoons \quad \text{OH}^- \quad \text{H}^+ \\
\text{Ammonium chloride completely dissociates:} \\
\text{NH}_4\text{Cl} \quad \text{(s)} \quad \rightarrow \quad \text{NH}_4^+ \quad \text{Cl}^- \\
\]

The ammonium ions form an equilibrium with the OH⁻ ions moving the water equilibrium to the right. Excess H⁺ are formed so pH is less than 7.
Alkaline Salt

In water you have:

\[ \text{H}_2\text{O}^{\text{(aq)}} \rightleftharpoons \text{OH}^-_{\text{(aq)}} + \text{H}^+ \]

Sodium ethanoate completely dissociates:

\[ \text{Na}^+\text{CH}_3\text{COO}^-_{\text{(aq)}} \rightarrow \text{Na}^+_{\text{(aq)}} + \text{CH}_3\text{COO}^-_{\text{(aq)}} \]

The ethanoate ions form an equilibrium with the H\(^+\) ions moving the water equilibrium to the right.

Excess OH\(^-\) ions are formed so the pH is greater than 7.

Calculating pH for a weak acid

Using the equilibrium constant (dissociation constant) from the reaction on page 47, the pH of a weak acid can be calculated:

\[ K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \]

Since \([\text{H}_3\text{O}^+] = [\text{A}^-] \):

\[ K_a = \frac{[\text{H}_3\text{O}^+]}{[\text{HA}]} \]
Taking logs of both sides:
\[ \log K_a = \log[H_3O^+]^2 - \log[HA] \]
\[ = 2\log[H_3O^+] - \log[HA] \]

but \(-\log[H_3O^+] = pH\), and assigning \(-\log K_a\) as \(pK_a\) we get:
\[-pK_a = -2pH - \log[HA] \]
\[pK_a = 2pH + \log[HA] \]

For a weak acid HA of concentration \(c\) mol l\(^{-1}\), \([HA]\) at equilibrium will be approximately equal to the original concentration \(c\) mol l\(^{-1}\), so:
\[pK_a = 2pH + \log c \]

Rearranging for \(pH\):
\[pH = \frac{1}{2}pK_a - \frac{1}{2}\log c \]

Questions

1. Calculate the \(pH\) of a 0.2 mol l\(^{-1}\) solution of ethanoic acid if \(K_a = 1.7 \times 10^{-5}\)

2. 0.02 mol l\(^{-1}\) benzoic acid \(C_6H_5COOH\), a monobasic acid, was found to have a \(pH\) of 2.94. Calculate the \(K_a\) of this weak acid.
Buffer solutions

A buffer solution is one in which the pH of the solution remains approximately constant when small amounts of acid or base are added or the solution is diluted with water.

An acidic buffer consists of a solution of a weak acid and one of its salts with a strong alkali. A basic buffer consists of a solution of a weak base and one of its salts with a strong acid. Both types of buffer solution work in the same way.

In an acid buffer solution, the weak acid supplies more hydrogen ions when the existing ones are removed by a base being added, while the salt of the weak acid provides the conjugate base to react with the hydrogen ions when small amounts of acid are added.

In a basic buffer solution, the weak base reacts with the hydrogen ions when acid is added and the salt provides the conjugate acid, which dissociates to replace the hydrogen ions when these are removed by addition of small amounts of base.

In consequence, the pH hardly changes in both cases. This principle is illustrated by an acid buffer of weak acid HA and the sodium salt of that acid NaA. In solution the following occurs:

\[
\text{HA}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-(aq) \quad \text{equilibrium}
\]

\[
\text{and} \quad \text{NaA}(aq) \rightarrow \text{Na}^+(aq) + \text{A}^-(aq) \quad \text{fully ionised}
\]

Addition of small volumes of H+ does not alter the pH greatly since an increase in [H+] favours the reverse reaction (HA formation) to maintain \( K_a \). Since [A\(^-\)] is high compared to [H\(^+\)] in the original buffer solution, its ability to remove H\(^+\) ions is substantial (but not infinite) and the pH of the solution is maintained.
In the same way, addition of $OH^-$ does not alter the pH greatly since the $OH^-$ ions combine with the $H^+$ of the weak acid until $K_w$ is attained. Removal of $H^+$ means that more HA ionises, giving $H^+ + A^-$ until equilibrium is re-established, and the pH of the solution is maintained:

$$K_a = \frac{[H_3O^+] [A^-]}{[HA]}$$

$$[H_3O^+] = \frac{[K_a][HA]}{[A^-]}$$

where $[HA]$ is the concentration of the acid as it is a weak acid and $[A^-]$ is the concentration of the salt as it is fully ionised and very little comes from the acid dissociating. Therefore:

$$[H_3O^+] = \frac{K_a[acid]}{[salt]}$$

If the buffer solution has water added to it, i.e. it is diluted, both $[acid]$ and $[salt]$ are equally affected and therefore $[H_3O+]$ and the pH remain unaltered.

An efficient buffer must have a reasonable reserve of HA and $A^-$ as the $H^+$ from HA removes added $OH^-$ and the $A^-$ removes added $H^+$. If $[acid] = [salt]$ the buffer will have the same ability to resist addition of $H^+$ and $OH^-$. Similar reasoning can be used to show how a basic buffer solution operates. The pH of an acid buffer solution can be calculated by conversion of the above equation,

$$pH = pK_a - \log \left( \frac{[acid]}{[salt]} \right)$$
Questions

1. Calculate the pH of the buffer solution made from 1.0 mol l⁻¹ methanoic acid and 1.78 mol l⁻¹ sodium methanoate solution. The pKₐ of methanoic acid is 3.8.

2. Calculate the pH of the buffer solution made from 0.1 mol l⁻¹ solution of ethanoic acid and potassium ethanoate. The pKₐ of ethanoic acid is 4.8.

Indicators

Indicators are used to determine the end-point in an acid-alkali titration. A suitable indicator must be chosen for any given reaction. Indicators are dyes whose colours are sensitive to pH. An indicator is usually a weak acid that dissociates as shown:

\[
\text{HIn(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{In}^-(aq)
\]

The unionised form of the indicator (HIn) has a different colour from its conjugate base (In⁻). The equilibrium constant (K_{In}) for the above equation is:

\[
K_{In} = \frac{\left[\text{H}_3\text{O}^+\right]\left[\text{In}^-\right]}{[\text{HIn}]}
\]
Rearranging this expression gives:

\[
\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{K_{\text{In}}}{[\text{H}_3\text{O}^+]} \]

This shows that the ratio of the two coloured forms is a function of the $H_3O^+$ concentration, i.e. the pH of the solution. It follows therefore that the colour of the indicator in any given solution depends on the relative concentrations (ratio) of the two coloured forms, which in turn is determined by the pH.

The theoretical point at which the colour change occurs is when $[\text{HIn}] = [\text{In}^-]$ and therefore $K_{\text{In}} = [\text{H}_3\text{O}^+]$. In other words, the colour change occurs when $pK_{\text{In}} = pH$. In practice the colour change cannot be seen when $[\text{HIn}] = [\text{In}^-]$ and it is only distinguishable when $[\text{HIn}]$ and $[\text{In}^-]$ differ by a factor of 10.

The pH range over which a colour change can be seen is therefore established using the expression:

\[
pH = pK_{\text{In}} \pm 1 \quad (\log 10 = 1)
\]

The appropriate indicator for any titration must be chosen such that the colour change occurs over the pH range when the pH is changing very rapidly. This means that the indicator must change colour during the addition of, say, half a drop of reagent. The indicator for a reaction can therefore be chosen by reference to titration curves, which are drawn from data obtained by measuring the pH of an acid continuously against the volume of alkali added during and after neutralisation.
**Strong Acid + Strong Alkali**

Note the rapid rise in pH as the end-point of the titration is approached. Thereafter the alkali added has nothing to react with and the pH continues to rise to a final value of 12. There is a region of rapid pH change around the end-point. For an indicator to be suitable, its pH range must fall within this region. It can be seen that for a strong acid and strong alkali a suitable indicator can have a colour change within the pH range 3-10.

**Strong Acid + Weak Alkali**

For a strong acid and weak alkali, a suitable indicator would change somewhere in the range 3-7.

**Weak Acid + Strong Alkali**

For a weak acid and strong alkali, a suitable indicator must change colour within the pH range 7-10.
Weak Acid + Weak Alkali

It is not possible to select any suitable indicator for a weak acid/weak alkali combination since the pH does not change rapidly enough at the end-point, i.e. the titration curve does not have an almost vertical section like the others.

The table below shows some common indicators, their pH range and colours.

<table>
<thead>
<tr>
<th>Name of indicator</th>
<th>pH of colour change</th>
<th>Colour (HIn)</th>
<th>Colour (In)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methyl orange</td>
<td>3.0–4.4</td>
<td>Orange</td>
<td>Yellow</td>
</tr>
<tr>
<td>Methyl red</td>
<td>4.2–6.3</td>
<td>Red</td>
<td>Yellow</td>
</tr>
<tr>
<td>Bromothymol blue</td>
<td>6.0–7.6</td>
<td>Yellow</td>
<td>Blue</td>
</tr>
<tr>
<td>Phenolphthalein</td>
<td>8.0–9.8</td>
<td>Colourless</td>
<td>Red</td>
</tr>
</tbody>
</table>

**Question**

From the table, calculate the pKa and Ka for the four indicators.