Equilibrium

Two-Way Reactions
STOP AND CHECK (PAGE 6)

- The reverse reaction is when the products (the right-hand side) react to form the original reactants (the left-hand side) again.

Equilibrium
STOP AND CHECK (PAGE 8)

- Equilibrium in a two-way reaction means that the forward reaction (making products from reactants) and the reverse reaction (making reactants from products) are both happening at the same rate or speed. This means that the amount of products you have and the amount of reactants you have doesn't change.
- The position of equilibrium refers to how much of the products are around, and how much of the reactants are around once equilibrium is hit. If there is more of one compared to the other, we say that side of the reaction is favoured.

Equilibrium Constant (K_c)
STOP AND CHECK (PAGE 9)

- To find the equilibrium constant, remember "products over reactants". Multiply all of your products on the top side of the fraction, and all of your reactants on the bottom side of the fraction. If a particular product or reactant has a big

number in front of it in the equation (i.e. the 3 in $3NO_2$), put that number as a power for that particular reactant (i.e. $[NO_2]^3$).

- When calculating the equilibrium constant, you do not include solids $_{(s)}$ and liquids $_{(l)}$. Things that are in an aqueous state $_{(aq)}$ and gases $_{(g)}$ should be included.

- $$\frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

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

The Value of the Equilibrium Constant (K_c)
STOP AND CHECK (PAGE 10)

- The numbers you use to calculate the equilibrium constant are the concentrations of the products and reactants.
- When K_c is 1, the concentration of products and reactants is exactly equal. If K_c is between 0.001 and 1000, it's close enough that they can be considered equal. A very small K_c (less than 0.001) tells us that there are a lot more reactants than products and the reverse reaction was favoured. Similarly, a very large K_c (more than 1000) tells us that there are a lot more products than reactants and the forward reaction was favoured.

Reaction Quotient (Q_c)
STOP AND CHECK (PAGE 18)

- K_c is the equilibrium constant, it's the proportion of products to reactants once equilibrium is reached. Q_c is the proportion of products over reactants before equilibrium has been reached, like if you've just started your reaction.
- If $Q_c = K_c$, you've reached equilibrium.
- If Q_c is smaller than K_c, then the forward reaction will dominate (it will happen faster than the reverse reaction) until equilibrium is reached.
- If Q_c is greater than Kc, then the reverse reaction will dominate (it will happen faster than the forward reaction) until equilibrium is reached.
- If $Q_c = K_c$, then the reaction has reached equilibrium.

Equilibrium

- Equilibrium is reached in a two-way chemical reaction when the rate at which reactants are reacting to form products is equal to the rate at which products are reacting to form reactants. This does not mean that the concentrations of products and reactants are equal, only that the rate at which they are reacting to form one another is the same (i.e. 5 reactants react to form products, and at the same time 5 products react to form reactants).
- By working out the reaction quotient (Q_c) and comparing it to the equilibrium constant (K_c) you can determine if the reaction has reached equilibrium. If the two values are the same, then the reaction has reached equilibrium.
- $K_c = \dfrac{[HI_{(g)}]^2}{[H_{(g)}][I_{(g)}]}$
    - $K_c = \dfrac{[0.43]^2}{[0.019][0.21]}$
    - $K_c = 46.3$
- $Q_c = \dfrac{[CO_{2(g)}][H_{2(g)}]}{[CO_{(g)}][H_2O_{(g)}]}$
    - $Q_c = \dfrac{[0.1][0.3]}{[0.3][0.5]}$
    - $Q_c = 0.2$
- Since $K_c = 1$ and $Q_c = 0.2$, Q_c is smaller than K_c. This means that the forward reaction will dominate in order to produce more products and increase Q_c until it becomes equal to K_c.

Disturbing Equilibrium

Le Chatelier's Principle

- Le Chatelier's principle states that "if an equilibrium is disturbed by changing the conditions, the position of the equilibrium moves to counteract the change". For example, if you remove some of the products in a reaction,

equilibrium will shift in favour of the forward reaction in order to replenish the products.

Change in Concentration
STOP AND CHECK (PAGE 14)

- If more ammonia (NH_3) was added, the reaction would shift to the left (reverse reaction) in order to remove some of the ammonia that was added and bring the reaction back to equilibrium.
- If more hydrogen (H_2) was added, the reaction would shift to the right (forward reaction) in order to use up some of the hydrogens that were added and bring the reaction back to equilibrium.

Change in Pressure
STOP AND CHECK (PAGE 17)

- Pressure is increased either by adding more particles to the system or by reducing the volume of the container.
- If the pressure was increased, the reaction would shift to the right (forward reaction) because that side produces fewer moles of gas (takes up less space) which lowers the pressure back down.
- If the pressure was decreased, the reaction would shift to the left (reverse reaction) because that side produces more moles of gas (takes up more space) which raises the pressure back up.

Change in Temperature
STOP AND CHECK (PAGE 19)

- Exothermic reactions release heat from the system into their surroundings (exo- like exit, so heat exits the system into the surrounding environment).
- Endothermic reactions absorb heat from their surroundings into the system (endo- like enter, so heat enters the system from its surroundings).
- If the forward reaction is exothermic, the reverse reaction must be endothermic. If the temperature is increased, the reverse (endothermic)

reaction which likes to absorb that heat will be favoured. The reverse reaction would make more N_2, so the concentration of N_2 would go up.

- If the temperature is decreased, then the forward reaction (exothermic) will be favoured in order to bring the temperature back up. The forward reaction would use up the N_2, so the concentration of N_2 would go down.

Catalysts
STOP AND CHECK (PAGE 20)

- Catalysts increase the rate of reaction, they do this for both the forward and reverse reactions equally.
- Since catalysts speed up both the forward and reverse reactions equally, equilibrium doesn't change. Since equilibrium doesn't shift, the concentration of N_2 stays the same.

Disturbing Equilibrium
QUICK QUESTIONS (PAGE 21)

- There are many factors that can cause a change in equilibrium. According to Le Chatelier's principle, equilibrium will shift to minimize the change to a system. This change could be from concentration, temperature or pressure.
- A change in the concentration of the products or reactants would cause a shift in equilibrium. If a concentration is increased equilibrium will shift to use up the substance that has been added. Similarly, if a concentration is decreased, equilibrium will shift to replenish the substance that has been lost. In both of these cases, the system acts to minimize change, as per Le Chatelier's principle.
- Temperature can also affect equilibrium. An increase in temperature will favour an endothermic reaction, and a decrease in temperature will favour an exothermic reaction. In the endothermic reaction, it will absorb the extra heat from the increase in temperature, thus minimizing the change in temperature. In the exothermic reaction, it will release heat and warm up the cooling environment, again acting to minimize the change to the system. In both of these cases the equilibrium shifts in order to minimize change as expected from Le Chatelier's principle.

- Changes in pressure can also cause a shift in equilibrium. An increase in pressure means less room for particles to move around, this favours the side of the reaction with fewer moles. Fewer particles for a given amount of space means pressure goes down. The system has acted to minimize the change as per Le Chatelier's principle. Similarly, with a decrease in pressure, the reaction will favour the side with more moles. More particles within a given amount of space mean an increase in pressure to counteract the initial decrease in pressure.

Solubility

What is Solubility?
STOP AND CHECK (PAGE 22)

- Solubility is on a spectrum, meaning that something could be completely dissolved, not at all dissolved, or anywhere in between (partially dissolved).
- The solubility of an ionic solid is a measure of how much of the solid will dissolve in the solvent. This is talking about the concentration of the solid once it's dissolved.
- Solubility is measured in moles per litre or $molL^{-1}$ and its symbol in equations is an s.

Solubility Constant (K_s)
STOP AND CHECK (PAGE 23)

- The solubility constant is just the proportion of products to reactants, only this time it's specifically for the reaction of dissolving an ionic solid.
- $K_s = [A^+_{(aq)}][B^-_{(aq)}]$
- For AB_2, it's $K_s = [A^+_{(aq)}][B^-_{(aq)}]^2$.
  - For A_2B, it's $K_s = [A^-_{(aq)}]^2[B^+_{(aq)}]$.
- $K_s = [Na^+_{(aq)}][Cl^-_{(aq)}]$
- $K_s = [Mg^{2+}_{(aq)}][Cl^-_{(aq)}]^2$
- $K_s = [K^+_{(aq)}][NO_3^-_{(aq)}]$

- $K_s = [Mg^{2+}_{(aq)}][OH^-_{(aq)}]^2$

Converting Between K_s and s

- You can calculate the solubility constant (K_s), substituting the solubility value for the concentrations in the K_s expression. Remember, to take into account if one of the products has a mole ratio other than 1:1 (i.e for A_2B, it will dissolve and form two A^- for every A_2B, so the molar ratio is 1:2, and you would use 2s instead of s).
- To find the solubility (s), substitute s for the concentrations in the K_s expression. Again, remember to take into account if any of the products have a mole ratio other than 1:1. Now rearrange for s and solve using your K_s value.

Precipitation

- Precipitation is when something comes out of a solution and forms a solid.
- A precipitate is a solid that's come out of solution, a solution is when that solid is dissolved.
- If your ionic product (Q_s) is greater than the solubility constant (K_s), a precipitate will form. If Q_s is smaller than K_s, then there will be no precipitate and more can be dissolved.

Common Ion Effect

- A common ion is an ion that is present in both the solute and the solution.
- Common ions decrease the solubility.
- Le Chatelier's principle tells us that, when the concentration of a product or reactant increases, the equilibrium responds by doing the same thing to the other side of the equation. In this case, the common ion increases the concentration of the dissolved product in the solution. The equilibrium

responds by creating more of the undissolved reactant. This decreases the solubility of the compound.

Solubility
QUICK QUESTIONS (PAGE 28)

- $K_s = [Ag^+_{(aq)}]^2[CO_3^{2-}_{(aq)}]$
    - $8.1\times10^{-12} = [2s]^2[s] = 4s^3$
    - $s = (\frac{8.1\times10^{-12}}{4})^{\frac{1}{3}} = 0.000127 \text{ molL}^{-1}$
- The Ag^+ from the AgCl already in solution would introduce the common ion effect. The added Ag^+ on the products side of the dissolving equation for Ag_2CO_3 would cause the system to favour the reverse reaction (forming a precipitate rather than dissolving) in order to use up the added Ag^+.
- $K_s = [Mg^{2+}_{(aq)}][OH^-_{(aq)}]^2$
    - $K_s = [s][2s]^2 = 4s^3$
    - $K_s = 4\times1.65\times10^{-4} = 2.87\times10^{-10}$
- To find the concentration of hydroxide ions:
    - $14 - pH = pOH$
    - $14 - 10 = 4$
    - $[OH^-] = 10^{-pOH}$
    - $[OH^-] = 10^{-4}$
    - $Q_s = [0.0075][10^{-4}]^2 = 7.5\times10^{-11}$
    - Since Q_s is smaller than K_s, no precipitate forms.

Acids and Bases

Acid and Base Definitions
STOP AND CHECK (PAGE 31)

- An acid is a proton donor, it donates a hydrogen atom. A base accepts protons, it receives a hydrogen atom.
- An amphiprotic substance is something that can either accept or donate a proton (hydrogen atom).

Acid-Base Reactions
STOP AND CHECK (PAGE 33)

- An acid is something that donates a proton.
- A base is something that accepts a proton.
- A conjugate acid or conjugate base is what is produced when an acid or base react. For an acid, its conjugate base is what's leftover when it donates a proton. For a base, its conjugate acid is what's produced when it receives a proton.

Acid and Base Strength
STOP AND CHECK (PAGE 35)

- A strong acid fully dissociates, a weak acid only partially dissociates.
- A strong base fully protonates, a weak base only partially protonates.
- Since in a weak acid or base reaction only some of the reactants have reacted to form products, the reverse reaction must also be happening. Since the reaction is going both ways it must be a two-way reaction.

The pH Scale
STOP AND CHECK (PAGE 37)

- The pH of a neutral solution is 7.
- Acids have a pH below 7, while bases have a pH above 7.
- $pH = -\log[H_3O^+]$

Self-Ionisation of Water
STOP AND CHECK (PAGE 38)

- Water self ionizes. Two water molecules, one acts as a base, the other as an acid, producing OH^- and H_3O^+.

- pH represents the relative concentrations of H_3O^+ and OH^-. If the concentration of H_3O^+ goes up, the pH will decrease. Likewise, if the concentration goes down, it will increase.
- pH represents the relative concentrations of H_3O^+ and OH^-. If the concentration of OH^- goes up, the pH will increase. Likewise, if the concentration goes down, it will decrease.

Calculating pH of Strong Acids
STOP AND CHECK (PAGE 40)

- You use the strong acid's concentration as the concentration of H_3O^+, so $-\log[acid]$.
- $-\log[0.01] = 2$
    - pH = 2
- $-\log[0.005] = 2.3$
    - pH = 2.3
- $-\log[0.2] = 0.70$
    - pH = 0.70

Calculating pH of Strong Bases
STOP AND CHECK (PAGE 42)

- The self ionization of water means that two water molecules can react together to form an H_3O^+ and OH^- ion.
- $K_w = [H_3O^+][OH^-]$
- Rearrange the above equation for OH^- or H_3O^+ and substitute 10^{-14} for K_w. So $[H_3O^+] = \frac{10^{-14}}{[OH^-]}$ and $[OH^-] = \frac{10^{-14}}{[H_3O^+]}$.
- pH is $-\log[H_3O^+]$, while pOH is $-\log[OH^-]$. It is very easy to convert between the two since pH + pOH = 14.
- $14 - -\log(0.01) = pH$
    - pH = 12
- $14 - -\log(2\times0.0025) = pH$
    - pH = 11.7

K_a and K_b
STOP AND CHECK (PAGE 46)

- The higher the K_a is, the stronger the acid, and the more it dissociates.
- The higher the K_b is, the stronger the base, and the more it protonates.

pH of Weak Acids
STOP AND CHECK (PAGE 48)

- There are two assumptions that you need to make when calculating the pH of a weak acid. The first is that the H_3O^+ present in the solution is only from the reaction of the weak acid. The second is that we assume that the final concentration of the weak acid (once it has reacted and reached equilibrium with the water) is the same as the initial concentration.
- The first assumption can be made because the contribution of H_3O^+ from the self ionization of water is very small and would not have any significant result on your calculation. The second assumption can be made since it is a weak acid and only a very small amount will react.

Acids and Bases
QUICK QUESTIONS (PAGE 55)

- $[OH^-] = (K_b \times [base])^{1/2}$
    - $[OH^-] = (1.8 \times 10^{-5} \times 0.3)^{1/2}$
    - $[OH^-] = 2.3 \times 10^{-3}$ molL^{-1}
    - $pOH = -\log[OH^-]$
    - $pOH = -\log[2.3 \times 10^{-3}] = 2.6$
    - $pH = 14 - pOH$
    - $pH = 11.4$
- $K_a = 10^{-pKa}$
    - $K_a = 10^{-7.46}$
    - $K_a = 3.47 \times 10^{-8}$
    - $[H_3O^+] = (K_a \times [acid])^{1/2}$
    - $[H_3O^+] = (3.47 \times 10^{-8} \times 0.005)^{1/2}$

- $[H_3O^+] = 1.3\times10^{-5}$ molL^{-1}
- pH = $-\log[H_3O^+]$
- pH = $-\log[1.3\times10^{-5}]$
- pH = 4.9

- $K_b = 10^{-pKb}$
  - $K_b = 10^{-3.37}$
  - $K_b = 4.27\times10^{-4}$
  - $[OH^-] = (Kb\times[base])^{\frac{1}{2}}$
  - $[OH^-] = (4.27\times10^{-4}\times0.01)^{\frac{1}{2}}$
  - $[OH^-] = 2.1\times10^{-3}$ molL^{-1}
  - pOH = $-\log[OH^-]$
  - pOH = $-\log[2.3\times10^{-3}]$
  - pOH = 2.7

- pH = 14 - pOH
  - pH = 11.3
  - pH of a strong acid is $-\log[acid]$.
  - pH of HNO_3 = $-\log[0.01]$
  - pH of NO_3 = 2
  - $[H_3O^+] = (K_a\times[acid])^{\frac{1}{2}}$
  - $[H_3O^+] = (1.8\times10^{-5}\times0.01)^{\frac{1}{2}}$
  - $[H_3O^+] = 4.2\times10^{-4}$ molL^{-1}
  - pH of CH_3COOH = $-\log[H_3O^+]$
  - pH of CH_3COOH = $-\log[4.2\times10^{-4}]$
  - pH of CH_3COOH = 3.4
  - The pH of the strong acid is lower than the pH of the weak acid for the same concentration.

Buffer Solutions

What Are Buffer Solutions?
STOP AND CHECK (PAGE 57)

- A buffer solution is made up of approximately equal amounts of a weak acid and its conjugate base, or a weak base and its conjugate acid.

The Function of Buffer Solutions

- Buffer solutions serve to counteract and minimize the change in pH when an acid or base is added to the solution.
- When a base is added, the weak acid in the buffer solution reacts with the OH^- to neutralize it and form water. Similarly, when an acid is added, the weak base in the buffer solution reacts with the H_3O^+ to neutralize it and form water.
- You can calculate the pH of a buffer solution using the formula:
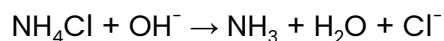
$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$$

Buffer Solutions

- Strong acids and strong bases can not create a buffer solution because they would fully react and neutralize each other. Only weak acids and weak bases can create a buffer solution because they have a two-way reaction, they do not fully react and neutralize each other. They stay in equilibrium until they can react with another strong acid or base to neutralize it.
- If a small amount of acid were added, the NH_3 would react to neutralize it:

$$NH_3 + H_3O^+ \rightarrow NH_4^+ + H_2O$$
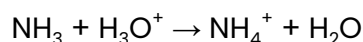
If a small amount of base were added, the NH_4Cl would react to neutralize it:

$$NH_4Cl + OH^- \rightarrow NH_3 + H_2O + Cl^-$$

- pH of a buffer solution = $pK_a + \log\left(\frac{[base]}{[acid]}\right)$
    - $pK_a = -\log K_a$
    - $pK_a = -\log 1.7 \times 10^{-5}$
    - $pK_a = 4.8$
    - $pH = 4.8 + \log\left(\frac{0.4}{0.45}\right)$
    - $pH = 4.7$

Relative Concentrations of Dissolved Species

Strong Acids and Bases
STOP AND CHECK (PAGE 63)

- The terms fully dissociate and fully protonate mean that all of the strong acid or all of the strong base has reacted.
- Hydroxide ions (OH^-) are present in acids because water has an equilibrium (two way) reaction with itself where it reacts to form H_3O^+ and OH^-. They are both products of this reaction, and the pH of a solution indicates which product is favoured but they are both always present.
- For strong acids, from highest concentration to lowest is:

$$[H_3O^+] > [acid] > [OH^-]$$

Likewise, for strong bases it is:

$$[OH^-] > [base] > [H_3O^+]$$

Weak Acids and Bases
STOP AND CHECK (PAGE 65)

- Strong acids fully dissociate, this means that almost all of the acid reacts with water to form H_3O^+. Weak acids only partially dissociate, so only a small portion of the weak acid reacts with water.
- Strong bases fully protonate, which means that almost all of the base reacts with water to form OH^-. Weak bases only partially protonate, so only a small portion of the weak base reactions with water.

Ionic Compounds (Salts)
STOP AND CHECK (PAGE 68)

- Ionic salts dissociate into their two ions in solution. In a neutral salt, there is no further reaction with water. In an acidic salt, one of the ions has acidic properties. The acidic ion will partially dissociate and undergo a secondary reaction with water to form H_3O^+ and its conjugate base. In an alkali salt, one of the ions has basic properties. The basic ion will partially dissociate and undergo a secondary reaction with water to form OH^- and its conjugate acid.
- One ion in the salt has acidic properties. It reacts with water and donates a proton to form H_3O^+.
- Neutral salts do not react with water to form either H_3O^+ or OH^-, so they do not alter the pH.
- $[Na^+] > [CH3COO-] > [OH^-] = [CH_3COOH] > [H_3O^+]$

Conductivity
STOP AND CHECK (PAGE 70)

- For a solution to be electrically conductive it must have ions in solution.
- Weak acids and bases are poor electrical conductors because they only partially dissociate, meaning few ions are formed. Few ions in a solution mean it will be poor electrical conductivity. Strong acids and bases on the other hand fully dissociate, meaning they form lots of ions. Lots of ions in solutions means there will be good electrical conductivity.
- Salts have a good electrical conductivity because initially, they dissociate into their ions when in solution. This results in a high concentration of ions in the solution, which gives them a good electrical conductivity.

Relative Concentration of Dissolved Species
QUICK QUESTIONS (PAGE 70)

- HNO_3 is a strong acid:
  $[H_3O^+] = [NO_3^-] > [OH^-]$
- NaOH is a strong base
  $[OH^-] = [Na^+] > [H_3O^+]$

- CH$_3$COOH is a weak acid

  $[CH_3COOH] > [H_3O^+] = [CH_3COO^-] > [OH^-]$

- NH$_3$ is a weak base

  $[NH_3] > [OH^-] = [NH_4^+] > [H_3O^+]$

- Na$_2$SO$_4$ is a neutral salt

  $[Na^+] > [SO_4^-] > [H_3O^+] = [OH^-]$

- CH$_3$COONa is an alkali salt

  $[Na^+] > [CH_3COO^-] > [CH_3COOH] = [OH^-] > [H_3O^+]$

- Ethanoic acid is a weak acid, meaning it only partially dissociates in water, so there is a lower concentration of charged particles present in the solution to carry the charge. Sodium ethanoate is a salt meaning it fully dissociates in water, so there is a large concentration of charged particles to carry the charge, so sodium ethanoate solution is more conductive than ethanoic acid solution.

- Sodium sulfate is a neutral salt, meaning it dissolves in water to form a neutral solution with a pH of 7. Sodium ethanoate is an alkali salt as the ethanoate ions produced when it dissolves react with water to produce hydroxide ions, so the solution is basic with a pH above 7. Both are salts which means they dissociate in water and while some of the ethanoate ions react with water to produce hydroxide ions, there is still about the same concentration of charged particles present in each solution so there is little difference in their conductivity.

- $K_b = 10^{-pKb}$
  - $K_b = 10^{-9.25}$
  - $K_b = 5.62\times10^{-10}$
  - $[OH^-] = (K_b\times[base])^{\frac{1}{2}}$
  - $[OH^-] = (5.62\times10^{-10}\times0.01)^{\frac{1}{2}}$
  - $[OH^-] = 2.37\times10^{-6} \: molL^{-1}$
  - $pOH = -log[OH^-]$
  - $pOH = -log[2.37\times10^{-6}]$
  - $pOH = 5.63$
  - $pH = 14 - pOH$
  - $pH = 14 - 5.63$
  - $pH = 8.37$ (3 s.f.)

Titration Curves

Titrations
STOP AND CHECK (PAGE 73)

- Titrations are done in order to determine the concentration of an unknown acid or base solution.
- The x-axis measures the volume of the known strong acid or base.
- The pH of the solution in your conical flask is measured on the y-axis. Initially, this is your unknown acid or base, but it gets gradually neutralized as you add your known acid or base.
- The three types of titrations you'll need to be familiar with are:
    - Strong acid + strong base.
    - Weak acid + strong base.
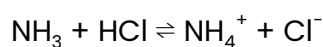   Weak base + strong acid.

Initial pH
STOP AND CHECK (PAGE 74)

- The initial pH is the pH of the acid or base before titration.

Equivalence Point
STOP AND CHECK (PAGE 75)

- Reaching the equivalence point tells us that all of the acid or base being titrated has been neutralized.
- NH_3 reacts with HCl in the following way:

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

This means that the titrated solution would contain NH_4^+ and Cl^-, as well as OH^- and H_3O^+.

The Half Equivalence Point

- Buffer solutions are formed by weak acids/bases and their conjugate base/acid. When a weak acid/base is titrated, it forms its conjugate base/acid, and so forms a buffer solution. Strong acids/bases do not have two-way reactions, so they do not form buffer solutions and so no buffer zone.
- The pH is equal to the pK_a of the weak acid at halfway to the equivalence point.

Indicators

- The role of indicators during titration is to indicate when the equivalence point has been reached.
- The most appropriate indicator to use is the one whose pK_a is closest to the pH at the equivalence point. Indicators change colour when the pH is equal to their pK_a.

Titration Curves

- Ethanoic acid is a weak acid, and sodium hydroxide is a strong base, so this is a weak acid/strong base titration.
- $[H_3O^+] = (K_a \times [acid])^{\frac{1}{2}}$
  - $[H_3O^+] = (1.7 \times 10^{-5} \times 0.01)^{\frac{1}{2}}$
  - $[H_3O^+] = 4.1 \times 10^{-4} \text{ molL}^{-1}$
  - $pH = -\log[H_3O^+]$
  - $pH = -\log[4.1 \times 10^{-4}]$
  - $pH = 3.39$
- $pK_a = -\log K_a$
  - $pK_a = -\log 1.7 \times 10^{-5}$

- $pK_a = 4.8$
  The buffer region occurs at a pH of 4.8, since at the buffer region the pH is equal to the pK_a of the weak acid.
- Initially, CH_3COOH is present, along with its conjugate CH_3COO^-. H_3O^+ and OH^- are also present. Once NaOH is added, Na^+ ions become present in the solution. As the NaOH is added, it reacts with the CH_3COOH. At the buffer region, roughly half of the CH_3COOH has reacted. As more NaOH is added, it eventually reacts with all of the CH_3COOH (at the equivalence point) leaving only CH_3COO^- and Na^+ along with H_3O^+ and OH^-.
- At the equivalence point n(acid) = n(base)
  - $n(CH_3COOH) = c(CH_3COOH) \times V(CH_3COOH)$
  - $n(CH_3COOH) = 0.0100mol/L \times 0.0500L$
  - $n(CH_3COOH) = 0.00050mol$
  - $V(NaOH) = \frac{n(NaOH)}{c(NaOH)}$
  - $V(NaOH) = \frac{0.0005}{0.0050}$
  - $V(NaOH) = 0.1L = 100mL$
  - 100mL of NaOH is needed to reach the equivalence point.
- The pH of the equivalence points will be greater than 7 as the conjugate base of the ethanoic acid reacts with the water to produce hydroxide ions, causing the solution to become basic.
  - $K_b = \frac{K_w}{K_a}$
  - $K_b = \frac{1 \times 10^{-14}}{1.7 \times 10^{-05}}$
  - $K_b = 5.88 \times 10^{-10}$
  - $K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$
  - $K_b = \frac{[OH^-]^2}{[CH_3COO^-]}$
  - $[OH] = (K_b \times [CH_3COO^-])^{\frac{1}{2}}$
  - $[OH] = (5.88 \times 10^{-10} \times 0.0100)^{\frac{1}{2}}$
  - $[OH] = 2.43 \times 10^{-6}$
  - $pOH = -log[OH^-]$
  - $pOH = -log(2.43 \times 10^{-6})$
  - $pOH = 5.62$
  - $pH = 14 - pOH$
  - $pH = 14 - 5.62$
  - $pH = 8.38$

- Thymol blue is the most suitable indicator as it changes colour at a pH of 8.9 which is the closest to the pH at the equivalence point.