Set 4. Equilibrium

Multiple Choice Questions

1. The reason for an increase in the rate of reaction at higher temperatures is due to:

   1) an increase in the number of collisions between the particles.
   2) lowering of the activation energy for the particles.
   3) an increase in the average kinetic energy of the reactant particles.

   (a) All are correct
   (b) 2 and 3 are correct
   (c) 1 and 3 are correct
   (d) 2 is correct
   (e) 1 is correct

2. Which of the following is **false**?

   (a) Activation energy is required for both exothermic and endothermic reactions.
   (b) Catalysts shift reaction equilibrium toward the side of the products.
   (c) Reaction rates depend on temperature, state of subdivision, concentration of the reactants and the presence of catalysts.
   (d) Enzymes are catalysts in living organisms.

3. A catalyst

   (a) increases the kinetic energy of the reaction.
   (b) provides a path of lower activation energy.
   (c) undergoes changes to speed up the rate of a reaction.
   (d) lowers the potential energy of the products compared to reactants.
   (e) increases the number of collisions of reactant molecules.

4. Use LCP to predict in which of the following the reaction will proceed more to the right by increasing the pressure.

   1) $2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g)$
   2) $2NO(g) \rightleftharpoons N_2(g) + O_2(g)$
   3) $N_2O_4(g) \rightleftharpoons 2NO_2(g)$
   4) $Ni(s) + 4CO(g) \rightleftharpoons Ni(CO)_4(g)$
   5) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

   (a) 1, 4, 5
   (b) 2, 3, 4
   (c) 1, 3
   (d) 2, 3
   (e) 1, 2, 3

5. Using Le Châtelier's Principle state which of the following reactions is the product formation favoured by decreased pressure.

1) $CO_2(g) + C(s) \rightleftharpoons 2CO(g)$ $\Delta H = +172$ kJ
2) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -91.8$ kJ
3) $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H = -21.7$ kJ
4) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ $\Delta H = +181$ kJ
5) $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$ $\Delta H = +484.6$ kJ

(a) 2, 3
(b) 3, 4
(c) 2, 4
(d) 1, 5
(e) 3

6. Using Le Châtelier's Principle state which of the following reactions is the product formation favoured by an increase in temperature.

1) $CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$ $\Delta H = -206.2$ kJ
2) $CO_2(g) + C(s) \rightleftharpoons 2CO(g)$ $\Delta H = +172.5$ kJ
3) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $\Delta H = -9.4$ kJ
4) $3O_2(g) \rightleftharpoons 2O_3(g)$ $\Delta H = +285$ kJ
5) $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$ $\Delta H = +484.6$ kJ

(a) 2, 4, 5
(b) 3, 5
(c) 1, 2, 5
(d) 1, 3
(e) 1, 3

7. Using Le Châtelier's Principle state which of the following reactions is the product formation favoured by a decrease in temperature.

1) $CO_2(g) + C(s) \rightleftharpoons 2CO(g)$ $\Delta H = +172.5$ kJ
2) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -91.8$ kJ
3) $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH$ $\Delta H = -21.7$ kJ
4) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ $\Delta H = +181$ kJ
5) $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$ $\Delta H = +484.6$ kJ

(a) 3
(b) 3 ,4
(c) 2, 3
(d) 1, 4, 5
(e) 2, 4

8. Using Le Châtelier's Principle state which of the following reactions is the product formation favoured by low pressure and high temperature.

1) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -91.8$ kJ
2) $2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$ $\Delta H = +566$ kJ
3) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $\Delta H = +9.4$ kJ
4) $3O_2(g) \rightleftharpoons 2O_3(g)$ $\Delta H = +285$ kJ
5) $N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g)$ $\Delta H = +66$kJ

(a) 5
(b) 4
(c) 1
(d) 2
(e) 3

9. Using Le Châtelier's Principle state which of the following reactions is product formation favoured by high pressure and low temperature.

1) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -91.8$ kJ
2) $2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$ $\Delta H = +566$ kJ
3) $3O_2(g) \rightleftharpoons 2O_3(g)$ $\Delta H = +285$ kJ
4) $H_2(g) + F_2(g) \rightleftharpoons 2HF(g)$ $\Delta H = -542$ kJ
5) $N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g)$ $\Delta H = +66.4$ kJ

(a) 4
(b) 5
(c) 2
(d) 1
(e) 3

10. Using Le Châtelier's Principle state which of the indicated changes will cause the reaction to proceed to the right.

1) $CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$ - add CH_4
2) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ - remove NH_3
3) $H_2(g) + F_2(g) \rightleftharpoons 2HF(g)$ - add F_2
4) $N_2(g) + 2O_2(g) \rightleftharpoons 2NO_2(g)$ - remove O_2
5) $BaO(s) + SO_3(g) \rightleftharpoons BaSO_4(s)$ - add BaO

(a) 3.5
(b) 2,3,5
(c) 1,4,5
(d) 1,4
(e) 2,3

11. Using Le Châtelier's Principle state which of the indicated changes will cause the reaction to proceed to the right.

1) $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$ - remove water
2) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ - add iodine
3) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ - add $CaCO_3$
4) $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ - remove O_2
5) $BaO(s) + SO_3(g) \rightleftharpoons BaSO_4(s)$ - add SO_3

(a) 1,3
(b) 1,4
(c) 3,4
(d) 2,3,5
(e) 2,5

12. Using Le Châtelier's Principle state what change will occur for the following reaction if a few drops of HCl are added.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

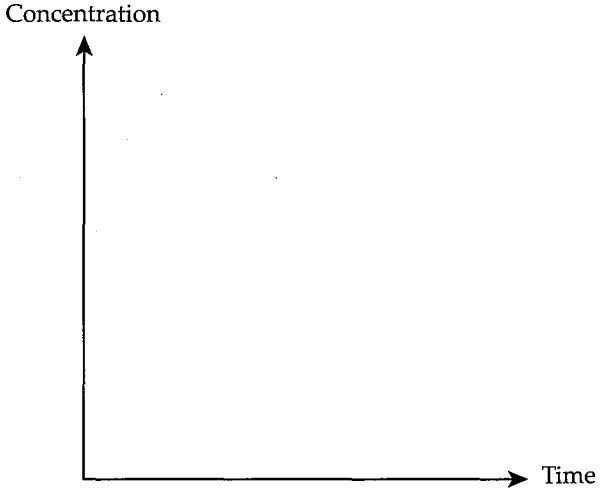
(a) A decrease in the number of acetate ions.
(b) A decrease in the number of hydronium ions
(c) An increase in the number of acid molecules ionised.
(d) An increase in the number of water molecules dissociated.

Longer Questions

1. If $Pb(NO_3)_2$ solution and KI solution are mixed in the correct stoichiometric ratio, a precipitate is formed.

(a) Write an equation for this reaction.

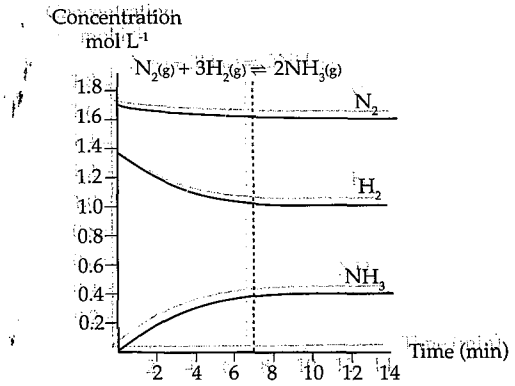
(b) Sketch a graph showing what happens to the concentration of Pb^{2+} ions, I^- ions and the precipitate during the course of this reaction.



(c) How do you know that the reaction has reached equilibrium?

(d) How could you show that the reaction has not stopped at equilibrium?

2. Study the equilibrium graph shown here.



(a) What substances are present at equilibrium?

(b) What is the final concentration of each of these substances?

(c) What is the concentration of each of these substances at time = 3 minutes?

(d) When did the concentration of H_2 reach 1.2 mol L^{-1}?

(e) When did the reaction reach equilibrium?

(f) What is the concentration of each substance at equilibrium?

3. (a) List ways in which a chemical reaction rate can be increased.

(b) In an aqueous solution of a mixture of $K_2Cr_2O_7$ and K_2CrO_4 in equilibrium, what would be the effect of adding the following:

$$Cr_2O_7^{2-}(aq) + 2OH^-(aq) \rightleftharpoons 2CrO_4^{2-}(aq) + H_2O(l)$$

(i) Sodium hydroxide solution?

(ii) Hydrochloric acid solution?

(iii) Water? _____

(c) When a bottle of soft drink is opened bubbles come out profusely, according to the equation:

$$CO_{2\,(aq)} \rightleftharpoons CO_{2\,(g)}$$

Why does this happen? Explain using Le Châtelier's Principle.

_____

_____

(d) When copper carbonate is heated, it decomposes into copper oxide and carbon dioxide. Explain both at molecular level and macroscopic level, what happens to the equilibrium position in this reaction if performed in:

(i) an open container

_____

_____

(ii) a sealed container

_____

_____

4. What effect would decreasing the pressure have on each of the following systems in equilibrium:

(a) $CO_{(g)} + 2H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$?

_____

_____

(b) $H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2HCl_{(g)}$?

_____

_____

(c) $2HgO_{(s)} \rightleftharpoons 2Hg_{(l)} + O_{2(g)}$?

_____

_____

(d) $2H_2O_{(g)} + 2I_{2(g)} \rightleftharpoons 4HI_{(g)} + O_{2(g)}$?

_____

_____

5. What effect would raising the reaction temperature have on the equilibrium position of each of the following equilibrium systems:

(a) $SO_{3(g)} + CO_{(g)} \rightleftharpoons SO_{2(g)} + CO_{2(g)}$? $\Delta H = -120\ kJ\ mol^{-1}$

_____

_____

_____

(b) $S_{(s)} + O_{2(g)} \rightleftharpoons SO_{2(g)}$? $\Delta H = -297 \text{ kJ mol}^{-1}$

(c) $COCl_{2(g)} \rightleftharpoons CO_{(g)} + Cl_{2(g)}$? $\Delta H = +108 \text{ kJ mol}^{-1}$

6. What would be the effect of lowering the temperature in the above systems?

(a) _____

(b) _____

(c) _____

7. Write the equilibrium expressions for the following reactions:

(a) $Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2OH^{-}_{(aq)}$ _____

(b) $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$ _____

(c) $2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$ _____

(d) $MnO_4^{-}{}_{(aq)} + 5Fe^{2+}_{(aq)} + 8H^{+}_{(aq)} \rightleftharpoons Mn^{2+}_{(aq)} + 5Fe^{3+}_{(aq)} + 4H_2O_{(l)}$

_____

(e) $2MnO_4^{-}{}_{(aq)} + 5H_2C_2O_{4(aq)} + 6H^{+}_{(aq)} \rightleftharpoons 2Mn^{2+}_{(aq)} + 10CO_{2(g)} + 8H_2O_{(l)}$

_____

8. A filter paper dipped in cobalt chloride solution (blue) can be used as an indicator of humidity in air (turns pink). The paper soaked in the solution and dried will absorb moisture from the air as follows:

$[CoCl_4]^{2-}_{(aq)} + 6H_2O_{(l)} \rightleftharpoons [Co(H_2O)_6]^{2+}_{(aq)} + 4Cl^{-}_{(aq)}$

Blue Pink

(a) Write an equilibrium expression for the reaction.

_____

(b) If you add some sodium chloride solution what will you observe? Explain your observation.

_____

(c) If you put these papers into a microwave oven and turned it on, what change will you observe to a blue cobalt chloride paper and to a pink cobalt chloride paper? Explain your observation.

_____

_____

_____

9. A copper chloride ionic equilibrium is set up for the following ions in 1 litre of solution:

$CuCl_2^-(aq) + Cl^-(aq) \rightleftharpoons CuCl_3^{2-}(aq)$ $\Delta H = -21$ kJ mol^{-1}

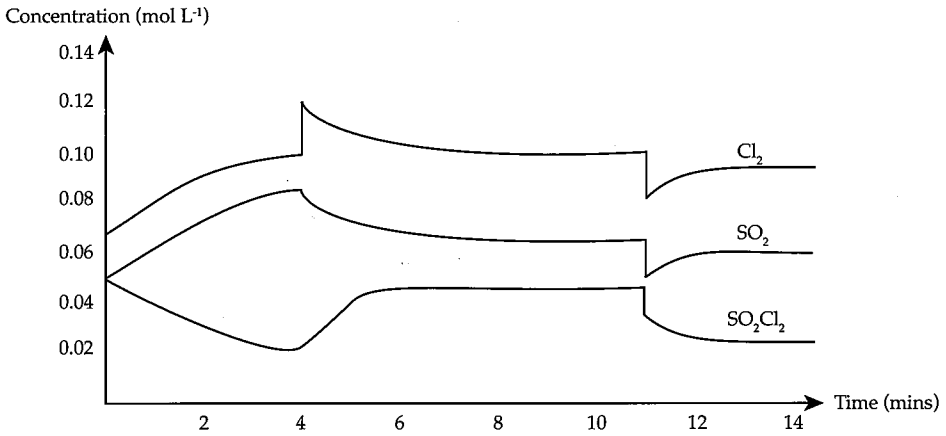
The pressure of the day was 100 kPa and the temperature was 25°C.

Write into the boxes below "increase", "decrease" or "no change" for the changes made.

Change made	Change in rate	Change in yield
Increase in pressure		
Increase in temperature		
Add some NaCl solid		
Divide the solution into 100 mL portions to increase the state of subdivision		

10. The graph below shows the results of the effects of changes to equilibrium in the reaction:

$SO_2Cl_2(g) \rightleftharpoons SO_2(g) + Cl_2(g)$ $\Delta H = +67$ kJ mol^{-1}



(a) What substances are present at the beginning of the reaction and what are their concentrations?

_____

(b) Write the equilibrium expression for this reaction.

_____

(c) Describe the change introduced and the effects that followed at time, t = 4 min.

(d) When did the system next reach equilibrium?

(e) What changes and effects occurred after time, t = 11 minutes?

11. Using Le Châtelier's Principle state the observation, direction of shift in equilibrium and an explanation for what happens in the following reactions, after each change is introduced:

(a) $Ba(OH)_{2(s)} \rightleftharpoons Ba^{2+}_{(aq)} + 2OH^-_{(aq)}$

White Colourless

(i) $BaCl_2$ solution is added to the mixture.

(ii) More $Ba(OH)_2$ solid is added to the mixture.

(b) $CaCO_{3(s)} + \text{heat} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

(i) A few drops of NaOH solution are added to the mixture.

(ii) Temperature of the system is decreased.

(iii) Water is added to the system.

(iv) Volume of the system is decreased.

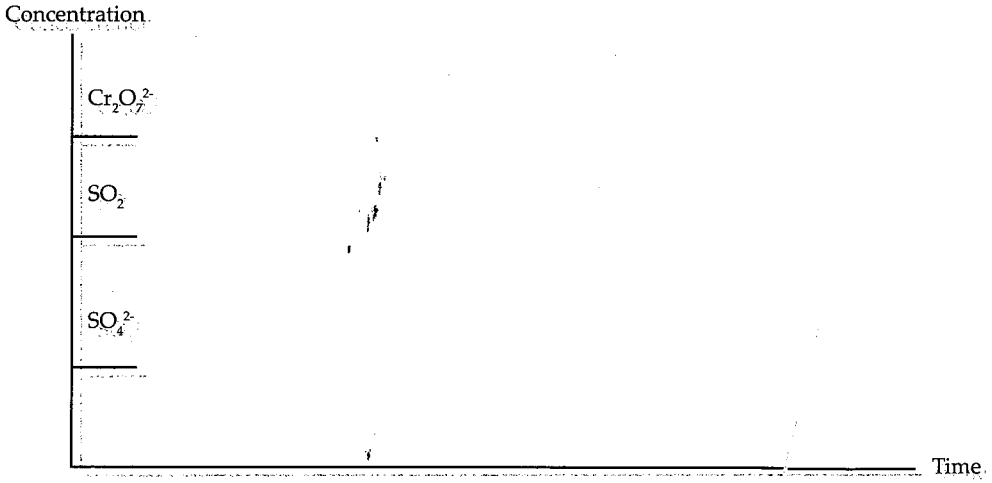
(c) $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)} + 99\ kJ$

(i) The volume of the system is increased.

(ii) An inert gas introduced in the system.

12. For the following reaction at equilibrium, changes are made to the reaction mixture as listed one after the other. Describe the effects that occur as a result of each of these changes. Show these on an equilibrium graph.

$Cr_2O_7^{2-}(aq) + 2H^+(aq) + 3SO_2(g) \rightleftharpoons 2Cr^{3+}(aq) + H_2O(l) + 3SO_4^{2-}(aq)$



(a) A solution of hydrochloric acid is added to the reaction mixture.

(b) A solution of $Ba(NO_3)_2$ is added to the mixture.

(c) The pressure on the system was increased.

(d) A solution of potassium hydroxide is added to the system.

(e) More SO_2 is pumped into the system.

13. Yellow bismuth oxybromide reacts with acid and makes the following equilibrium:

$BiOBr_{(s)} + 2H^+_{(aq)} \rightleftharpoons Bi^{3+}_{(aq)} + Br^-_{(aq)} + 2H_2O_{(l)}$

Different test tubes a, b and c containing BiOBr and HCl were prepared and subjected to 3 different experiments shown below as changes.

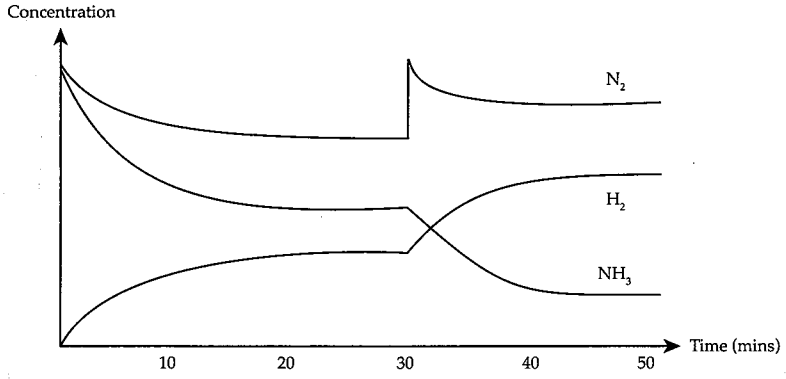
Using collision theory and, using Le Châtelier's Principle, complete the table below to show how the equilibrium will shift with these changes and explain the reason.

Tube No.	Change imposed	Shift in equilibrium Left, Right or None	Explanation
a	5 mL of water added		
b	A few drops of HBr added		
c	A few drops of $AgNO_{3\,(aq)}$ added		

14. Consider the ammonia production equation below:

$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g) \quad \Delta H = -92\ kJ\ mol^{-1}$

The graph below shows the concentrations of the three gases involved in the reaction:



Explain:

(i) Why do the pressures of all gases stabilise around the 20 minute mark?

_____

(ii) Why the partial pressure of the H_2 decreases more rapidly than that of the N_2?

_____

(iii) What has occurred at the 30-minute mark to cause the changes shown in the graph?

_____

(iv) What will the change made at the 30-minute mark have made by the time it has reached 50 minutes to the rate of:

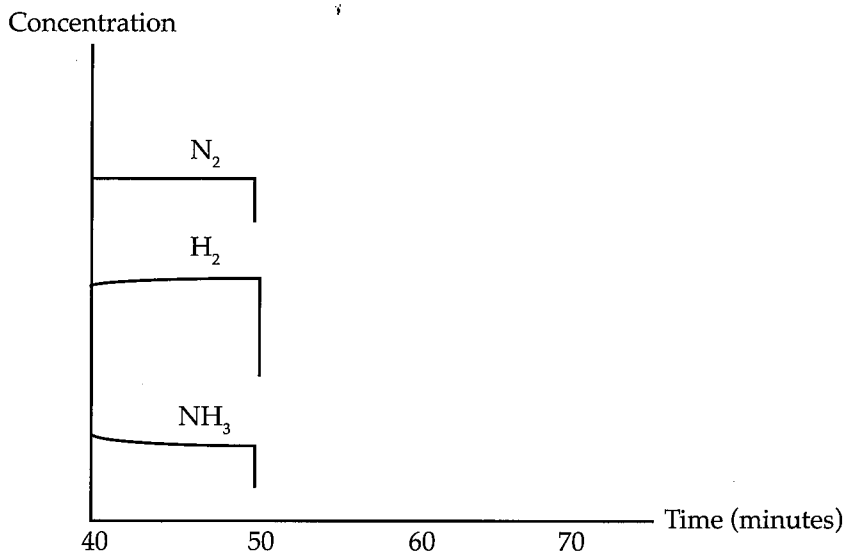
The forward reaction? _____

The reverse reaction? _____

(v) Use the Collision Theory to explain your answers to part (iv).

_____

At 50 minutes, the contents of the reaction vessel are rapidly expanded by increasing the volume. The changes in the concentrations of the gases are shown on the following graph.



(vi) Complete the graph above by drawing over it to show how the concentration of each gas will change as it moves to reach equilibrium at a time of 70 minutes.

15. The amount of gaseous atmospheric CO_2 is increasing as a direct result of human activities such as deforestation, cement manufacture and the burning of fossil fuels. Atmospheric CO_2 absorbs heat which otherwise would have radiated out from the planet through the atmosphere, resulting in global warming. Increased atmospheric CO_2 has also increased the levels of aqueous CO_2 dissolving into the oceans and ocean acidity.

(a) Write an equation showing the equilibrium that occurs between gaseous atmospheric CO_2 and aqueous oceanic CO_2 concentrations.

_____

(b) Given that CO_2 dissolving into the oceans is an exothermic process, what effect could increasing oceanic temperatures have on the equilibrium equation in (a) above?

_____

(c) Write an equilibrium equation showing how the increased dissolution of CO_2 into the ocean increases the acidity.

(d) Explain, using collision theory and with reference to the relative rates of the forward and reverse reactions from your answer to (c) above, how the increase of aqueous oceanic CO_2 contributes to ocean acidification.

Increased acidity removes CO_3^{2-} from the water, reducing the ability of sea organisms to build their calcium carbonate shells and body structures. In addition the increased acidity can cause their shells to dissolve.

(e) Using your equation from (c) above, and Le Châtelier's Principle, explain how the concentration of CO_3^{2-} in sea water is reduced.

(f) Give an ionic equation for the dissolution of calcium carbonate caused by increased ocean acidity.

products will increase. Rate will increase but no equilibrium shift.

9.

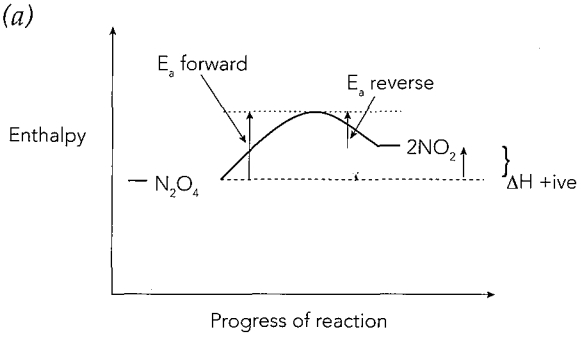
(i) *Pressure will increase as there are more particles. Yield and rate will remain constant as particle spacing remains the same.*

(ii) *For pressure to remain constant volume must increase therefore both rates decrease, but yield will increase as equilibrium shifts right.*

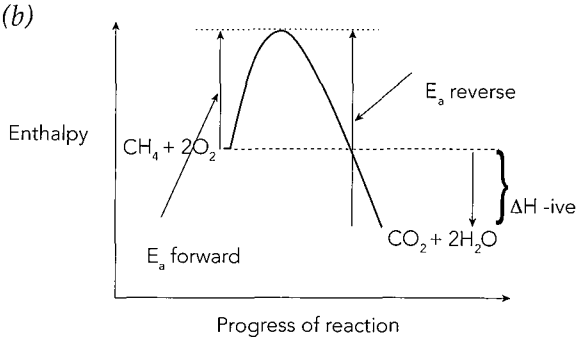
10. *Pressure will be reduced as volume increases. Yield will reduce as equilibrium shifts left and so $[SO_2]$ will increase but up to a lower level than before – hence rate will be lower.*

11.

(a)



(b)



Set 3 Changing Temperatures

1. *Pressure will decrease but equilibrium shifts to the right as exothermic and yield must increase. Forward rate will decrease as temperature is reduced and equilibrium shifts to the right because reverse rate decreases more than the forward rate.*
2. *Pressure must increase with temperature. Yield must decrease for exothermic reactions as equilibrium shifts to the left. Rate must increase with higher temperature.*
3.

(i) *Pressure must increase with temperature as K increases for endothermic reactions ($K = [CO_2]$).*
Yield will increase as equilibrium shifts to the right. Rate always increases with temperature.

(ii) *Mass of $CaCO_3$ will be reduced as equilibrium shifts to the right.*

(iii) *$K = [CO_2]$.*

4.

(i) *Pressure must increase with temperature. Yield must increase with temperature for the endothermic reaction. Rate must increase with temperature.*

(ii) *More iodine would be produced so colour becomes darker.*

5. *Pressure must increase with temperature. Yield must increase with temperature for an endothermic reaction. Rate must increase with temperature.*
6.

(i) *Exothermic equilibrium shifts to the left so mass of AgCl decreases. Forward rate must increase with higher temperature but reverse rate increases more.*

(ii) *Rate of forward reaction would decrease as the equilibrium shifts to the left. AgCl decreases.*

7.

(i) *Pressure must increase with temperature. Yield must decrease with temperature for an exothermic reaction. Forward rate must increase with higher temperature but reverse rate increases more so the equilibrium shifts to the left.*

(ii) *Mass and volume of bromine would increase.*

8.

(i) *Pressure must increase with temperature. Yield must increase with temperature for an endothermic reaction. Rate must increase with higher temperature.*

(ii) *Mass of ZnO would decrease as the equilibrium shifts to the right.*

9. *Pressure must increase with temperature. Yield must increase with temperature for an endothermic reaction. Rate must increase with higher temperature.*
10.

(i) *$[Cl_2]$ would decrease as the equilibrium shifts to the left.*

(ii) *$FeCl_2$ is a solid and so its concentration cannot change but mass would increase.*

(iii) *The mass of iron would decrease as the equilibrium shifts to the right.*

(iv) *Pressure must decrease with a temperature drop as Cl_2 molecules collide less.*

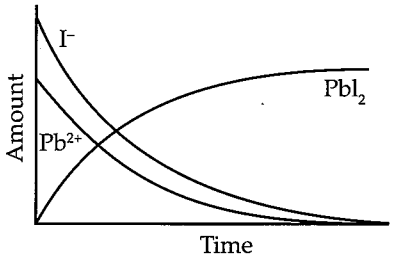
(v) *Total enthalpy change would be -342 + -57 = -399 kJ.*

Set 4 Equilibrium

1. c, 2. b, 3. b, 4. a, 5. d, 6.a, 7. c, 8. d, 9. d, 10. e, 11. e, 12. a

swers to longer questions

(i) $Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$



The amount of the precipitate formed, its colour or intensity of the yellow colour, or any other observable changes have ceased. This could only be shown using tagged radioactive isotope mixed with normal iodine. It could be shown that the proportion of the radioactive isotopic iodine continues changing during the equilibrium even though the amounts remain constant.

N_2, H_2, and NH_3 are present at equilibrium.
Final concentrations: $[N_2] = 1.60$ mol L^{-1}, $[H_2] = 1.00$ mol L^{-1}, $[NH_3] = 0.40$ mol L^{-1}
Concentrations after 3 minutes: $[N_2] = 1.65$ mol L^{-1}, $[H_2] = 1.10$ mol L^{-1}, $[NH_3] = 0.25$ mol L^{-1}
About 2 minutes after the reaction commences.
At the seventh minute after the reaction commences.
Same as given for b) above.

Increase the temperature of the reaction mixture.
Increase the pressure in gaseous reactions.
Increase the concentrations of the reactants.
Increase the surface area of any solids which are reacting.
Use a catalyst
$Cr_2O_7^{2-}(aq) + H_2O(l) \rightleftharpoons 2CrO_4^{2-}(aq) + 2H^+(aq)$
(i) Reaction would move to the right and colour would turn more yellow.
(ii) Reaction would move to the left and colour would turn more orange.
(iii) All ion concentrations would become less initially, then by LCP reaction would move to the right to increase the concentration of ions.
When the bottle is opened CO_2 gas is lost from the left side of the equation, so, by LCP, the reaction will move right to replace it.
(i) In an open container the CO_2 is lost and so by LCP reaction moves right to produce more. Eventually all the $CaCO_3$ will have decomposed.
(ii) In a sealed container the CO_2 cannot escape and an equilibrium is reached where the forward rate = reverse rate and the mass of $CaCO_3$ will become constant.

4.
(a) A decrease in pressure will shift the equilibrium to the left, forming more of the reactants from the products.
(b) A decrease in pressure has no effect because the number of gaseous moles are equal on both sides.
(c) A decrease in pressure will shift the equilibrium to the right (towards greater number of gaseous moles), forming more products from the reactants.
(d) A decrease in pressure will shift the equilibrium to the right (towards a greater number of gaseous moles), forming more products.

5.
(a) Raising the temperature will shift the equilibrium to the left, forming more reactants as this is an exothermic reaction.
(b) Raising the temperature will shift the equilibrium to the left, forming more reactants as this is an exothermic reaction.
(c) Raising the temperature will shift the equilibrium to the right, forming more products as this is an endothermic reaction.

6. The opposite effect to what is stated in question 5 will occur in each case.

7.
(a) $K = [Ca^{2+}][OH^-]^2$

(b) $K = \dfrac{[NH_3]^2}{[N_2][H]^3}$

(c) $K = \dfrac{[NO_2]^2}{[NO]^2[O_2]}$

(d) $K = \dfrac{[Mn^{2+}]^2[Fe^{3+}]^5}{[MnO_4^-]^2[Fe^{2+}]^5[H^+]^6}$

(e) $K = \dfrac{[Mn^{2+}]^2[CO_2]^{10}}{[MnO_4^-]^2[H_2C_2O_4]^5[H^+]^6}$

8.
(a) $K = \dfrac{[Co(H_2O)_6]^{2+}[Cl^-]^4}{[CoCl_4^-]^2}$

(b) If you sprinkle some NaCl solution, the increased concentration of chloride ions will shift the equilibrium to the left and the solution will become blue.
(c) After microwaving the papers the blue paper will remain blue and the pink one will turn blue. The shift in equilibrium is to the left as water is removed from the paper.

9.

Change made	Change in rate	Change in yield
Increase in pressure	No change	No Change
Increase in temperature	Increase	Decrease
Add some NaCl solid	Increase	Increase
Divide the solution into 100 mL portions to increase the state of subdivision	No change	No change (not changing concentrations)

10.

(a) *At the beginning of the reaction, SO_2, Cl_2, and SO_2Cl_2 are all present. ($SO_2 = 0.05$ M, $Cl_2 = 0.068$M, $SO_2Cl_2 = 0.05$ M)*

(b) $K = \frac{[SO_2][Cl_2]}{[SO_2Cl_2]}$

(c) *Chlorine was pumped into the system. The increased concentration of one of the products shifts the equilibrium to the left. Chlorine starts reacting with SO_2, producing SO_2Cl_2. Therefore $[SO_2]$ begins to decrease and $[SO_2Cl_2]$ begins to increase.*

(d) *At t = 9 mins.*

(e) *Volume of vessel is reduced so all concentrations decrease.*

11.

(a) *i) After $BaCl_2$ solution is added, the concentration of Ba^{2+} ions increases. Equilibrium shifts to the left. More $Ba(OH)_2$ is produced. The solution becomes cloudy initially, then becoming whiter.*
ii) Addition of $Ba(OH)_2$ solid does not eventually produce any effect as K must stay constant.

(b) *i) NaOH solution reacts and will dissolve acidic CO_2 gas. The equilibrium will shift to the right to produce more CO_2. More $CaCO_3$ will decompose.*
ii) This is an endothermic reaction. Decrease in temperature will shift the equilibrium to the left, forming more $CaCO_3$. CO_2 gas is reduced because K reduces.
iii) The reaction moving to the right to produce some more CO_2, as some of this gas dissolves in the added water.
iv) A decrease in pressure will drive the reaction to the side of more number of gaseous moles. Equilibrium will shift to the right. More CO_2 will be produced but the concentration will remain the same so that K can stay the same. $K = [CO_2]$. $CaCO_3$ decreases.

(c) *i) When the volume of the system is increased, the reaction moves in the direction of greater number of gaseous moles (left). More reactants are produced and the system gets cooler as a result.*
ii) Introduction of an inert gas does not change the concentration of the gases and so K remains the same and there is no change in yield or rate.

12. $Cr_2O_7^{2-}(aq) + 2H^+(aq) + 3SO_2(g)$ *Orange* $\rightleftharpoons 2Cr^{3+}(aq) + H_2O(l) + 3SO_4^{2-}(aq)$ *Green*

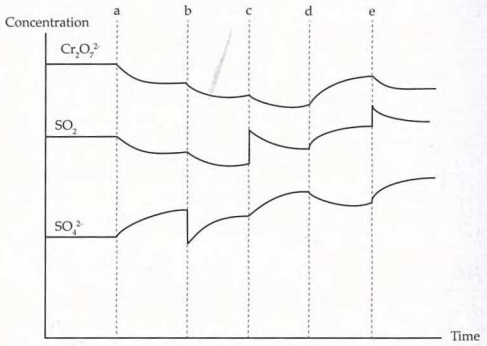
(a) *When a solution of HCl is added to the mixture, $[H^+]$ increases. The equilibrium shifts to the right. More Cr^{3+} and SO_4^{2-} are produced and the mixture becomes greener..*

(b) *The added Ba^{2+} ions will react with SO_4^{2-} ions to produce $BaSO_4(s)$ and decrease its concentration. Equilibrium will shift to the right in order to produce more SO_4^{2-} ion. The mixture becomes more greenish.*

(c) *The reaction will then move to the right according to LCP to partially increase the concentrations again and so the colour will become greener.*

(d) *The added OH^- ions will react with the H^+ ions in the mixture to produce H_2O, thus decreasing its concentration. Equilibrium will shift to the left producing more reactants. The solution will become more orange.*

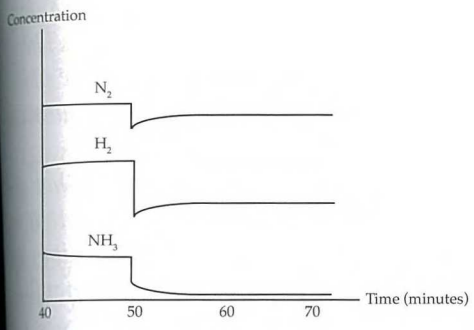
(e) *Increased concentration of the reactant SO_2 will shift the equilibrium to the right leading to the production of more Cr^{3+} and the other products. The solution will become greener.*



13.

Tube No.	Change imposed	Shift in equilibrium	Explanation
A	5 mL of water added	None	All ions are diluted but no increase. Both rates reduced.
B	A few drops of Br added	Left	Concentration of Br^- increases so reaction moves left by LCP to decrease Br^- concentration
C	A few drops of $AgNO_{3(aq)}$ added	Right	Ag^+ ion reacts with the Br^- ion to form a precipitate of AgBr and removes Br^-. So reaction moves right by LCP to increase Br^- concentration.

The reaction has reached equilibrium.
For every N_2 molecule used in the reaction 3 H_2 molecules are use up so the H_2 will decrease 3 times faster than the N_2.
More N_2 was introduced into the vessel.
Forward reaction rate and reverse reaction rate will have increased but forward rate increases more than the reverse rate.
With more nitrogen present, the particles are closer and so there will be more collisions per second and the forward rate will go up producing more product. With more product, the reverse collision rate will then rise.



$CO_2(g) \rightleftharpoons CO_2(aq)$
Increasing oceanic temperatures would cause the equilibrium to shift in the reverse direction so more $CO_2(g)$ would be produced.
$CO_2(aq) + H_2O(l) \overset{1}{\rightleftharpoons} H^+(aq) + HCO_3^-(aq) \overset{2}{\rightleftharpoons} 2H^+(aq) + CO_3^{2-}(aq)$
As carbon dioxide dissolves into the ocean LCP predicts that equilibrium 1 shifts in the forward direction to partially counteract the imposed change and the concentration of hydrogen ions and hydrogencarbonate ions increases. In turn, LCP predicts that equilibrium 2 will shift in the forward direction, further increasing the concentration of hydrogen ions. The pH decreases and acidity increases.
Increasing the CO_2 concentration increases the collisions between CO_2 and H_2O molecules and so the rate of the forward reaction 1 increases relative to the rate of its reverse reaction. This in turn increases the collisions between hydrogen ions and hydrogencarbonate ions and so the rate of the forward reaction 2 increases relative to the rate of its reverse reaction. More hydrogen ions contribute to increased acidity.
Increasing hydrogen ion concentration causes the equilibrium 2 above to shift in the reverse direction in order to partially counteract the imposed change. This causes the carbonate ion concentration to decrease.
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

Chapter 2. Acids and Bases

Set 1 Brønsted-Lowry Acids and Bases

1.

(a)

(i) $2HCl(aq) + Mg(s) \rightarrow MgCl_2(aq) + H_2(g)$
$2H^+(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + H_2(g)$

(ii) $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(iii) $2HNO_3(aq) + CaO(s) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$
$2H^+(aq) + CaO(s) \rightarrow Ca^{2+}(aq) + H_2O(l)$

(iv) $2HBr(aq) + K_2CO_3(aq) \rightarrow 2KBr(aq) + H_2O(l) + CO_2(g)$
$2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

(b) *Davy identified acids as substances that contain hydrogen that could be replaced by metals. In equation (i) and (iii) the metal has replaced the hydrogen in the acid to produce the salt.*

(c) *Magnesium metal is not considered a 'Davy' base because when it reacts with an acid it does produce a salt, but not water. In our terms, the reaction is not a neutralisation.*

2.

(a) *Weak Arrhenius acid, since it does produce hydrogen ions in water but does not ionise completely.*
$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq)$

(b) *Weak Arrhenius base, since it does produce hydroxide ions in water but does not onise completely.*
$NH_3(g) + H_2O(l) \rightleftharpoons + NH_4^+(aq) + OH^-(aq)$

(c) *Weak Arrhenius acid, since it does produce hydrogen ions in water but does not ionise completely.*
$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$

(d) *Strong Arrhenius base, since it dissociates completely to produce OH^- ions in solution.*
$LiOH(s) \rightarrow Li^+(aq) + OH^-(aq)$

(e) *Strong Arrhenius acid, since its first ionisation occurs completely.*
$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$

3. *(i) Acid = H_2O, Base is CN^-*
(ii) Acid = CH_3COOH, Base = S^{2-}
(iii) Acid = HS^-, Base = CO_3^{2-}

4. *(i) H_2O/ OH^-, HCN/ CN^-*
(ii) CH_3COOH/ CH_3COO^-,
(iii) HS^-/ S^{2-}, HCO_3^-/ CO_3^{2-}

5. *Theoretically the HCl solution should give the same number of H^+ ions as the H_2SO_4 as the latter is diprotic and the acids are both strong. However, not both of the hydrogens from the H_2SO_4 are fully ionised in solution. One is a strong ionisation and the second is weak so for HCl, moles of H^+ is $0.2 \times 1 = 0.2$*