Practice Questions

1 Write the equilibrium constant expression for each of the following.

Equation	$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$
Equilibrium constant expression	

(1 mark)

Equation	$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,Cl^-(aq)$
Equilibrium constant expression	

(1 mark)

2. The uptake of carbon dioxide from the atmosphere by the oceans is leading to gradual acidification of the oceans (i.e. the oceans are becoming less alkaline). When carbon dioxide dissolves, it reacts with water to form carbonic acid, which in turn forms hydrogencarbonate and then carbonate ions.

(a) Write equilibrium equations that show the formation of these products in water.

One of the most significant consequences of ocean acidification is the effect on shellfish and other marine life that produce and rely on calcium carbonate as a major component of the exoskeleton or other supporting structure. If the water is sufficiently acidic, the carbonate structures may not form completely. Ocean acidification is thought to lead to a reduction in the availability of carbonate ions. Further reaction of the dissolved carbon dioxide occurs as shown below.

$$CO_2(g) + CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons 2\,HCO_3^-(aq)$$

(b) What can you conclude about the magnitude of the equilibrium constant for the above reaction, and the relative proportions of products and reactants in the system?

(2 marks)

3.

(a) Write the equilibrium constant expression for each of the following.

Equation	$Co(H_2O)_6^{2+}$ (aq) + $4Cl^{-1}$ (aq) $\rightleftharpoons$ $CoCl_4^{-2}$ (aq) + $6H_2O$ (l)
Equilibrium constant expression	

(1 mark)

Equation	$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$
Equilibrium constant expression	

(1 mark)

A gaseous chemical system was observed to be at equilibrium at room temperature. The colour of the system was pink due to the presence of a greater concentration of red reactant A molecules and the low concentration of blue coloured B products. The system was then heated to 100°C and a distinctly blue colour appeared in the reaction flask. This blue colour reverted back to the original pink colour when the heat source was removed and room temperature was reattained.

A (g)(red) $\rightleftharpoons$ B (g) (blue)

(b) What can you conclude about the size of the equilibrium constant at room temperature and at 100°C? Explain. (2 marks)

(c) State Le Chateliers principle and use it to determine if the forward reaction (as written above) is exothermic or endothermic.

(2 marks)

87 Consider the following equations which show reversible reactions.

Reaction 1 $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$ $\Delta H^{\ominus} = -900$ kJ

Reaction 2 $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H^{\ominus} = -91$ kJ

a) In industry these reactions are carried out in the presence of catalysts. A platinum catalyst is used in *Reaction 1* and copper catalyst is used in *Reaction 2*.

State and explain the effect on the yield of a reaction when a catalyst is used. (2 marks)

4.

b) State and explain which of the above reactions will give an increase in the yield of product(s) when the pressure is increased at a constant temperature. (3 marks)

c) State and explain the effect on the yield of NO(g) when the temperature is increased in *Reaction 1* at a constant pressure. (3 marks)

d) i) Write an expression for the equilibrium constant, K_c, for *Reaction 2*. State the units of K_c. (2 marks)

5.

Chlorine for disinfection

Chlorine is used in water treatment for disinfection.

When chlorine is added to pure water, hypochlorous acid (HOCl) and hydrochloric acid (HCl) are formed.

$Cl_2(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + H^+(aq) + Cl^-(aq)$

The principal disinfecting action of aqueous chlorine is due to the hypochlorous acid formed.

Hypochlorous acid dissociated into hydrogen ions and hypochlorite ions.

$HOCl(aq) \rightleftharpoons H^+(aq) + OCl^-(aq)$

The concentrations of hypochlorous acid and hypochlorite ions in chlorinated water depend on the pH of the water.

Instead of using chlorine gas, some plants apply sodium hypochlorite or calcium hypochlorite to water. Sodium hypochlorite completely dissociates in water to form sodium ions and hypochlorite ions. In solution, the hypochlorite ions hydrolyze to form the disinfectant hypochlorous acid according to the following equation:

$OCl^-(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + OH^-(aq)$

a) State Le Chatelier's principle. (2 marks)

b) Use Le Chatelier's principle to explain how the pH of the chlorinated water will affect the concentrations of hypochlorous acid and hypochlorite ions in the water. (4 marks)

c) Write the equilibrium constant for the following reaction i .

$OCl^-(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + OH^-(aq)$

What is the equilibrium constant K_2 for the following reaction?

$2HOCl(aq) + 2OH^-(aq) \rightleftharpoons 2OCl^-(aq) + 2H_2O(l)$

d) The hypochlorous acid produced in a solution of sodium hypochlorite can react further to produce small amount of chlorine according to the following equation:

$HOCl(aq) + H^+(aq) + Cl^-(aq) \rightleftharpoons Cl_2(g) + H_2O(l)$

What will happen to the concentration of chlorine if a little sodium hydroxide solution is added to a sodium hypochlorite solution? Explain your answer. (4 marks)

Anskey

(a) Write the equilibrium constant expression for each of the following.

Equation	$Co(H_2O)_6^{2+}$ (aq) + $4Cl^{-1}$ (aq) $\rightleftharpoons$ $CoCl_4^{-2}$ (aq) + $6H_2O$ (l)
Equilibrium constant expression	$\frac{[CoCl_4^{-2}]}{[Co(H_2O)_6^{2+}][Cl^-]^4}$

(1 mark)

Equation	$CaCO_3$(s) $\rightleftharpoons$ CaO(s) + CO_2(g)
Equilibrium constant expression	$[CO_2]$

(1 mark)

A gaseous chemical system was observed to be at equilibrium at room temperature. The colour of the system was pink due to the presence of a greater concentration of red reactant A molecules and the low concentration of blue coloured B products. The system was then heated to 100°C and a distinctly blue colour appeared in the reaction flask. This blue colour reverted back to the original pink colour when the heat source was removed and room temperature was reattained.

A (g)(red) $\rightleftharpoons$ B (g) (blue)

(b) What can you conclude about the size of the equilibrium constant at room temperature and at 100°C? Explain. (2 marks)

At 100°C there is more products than at 25°C, so K at 100°C is larger than K at 25°C 1

(c) State Le Chateliers principle and use it to determine if the forward reaction (as written above) is exothermic or endothermic.

(2 marks)

Le Chateliers principle states that when an imposed change is applied to a closed system that is at equilibrium. The system will partially counteract the imposed change 1 to reestablish

Since heating would result in system favouring the endothermic reaction, to lower heat, the forward reaction is endothermic 1

87 Consider the following equations which show reversible reactions.

Reaction 1 $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$ $\Delta H^{\ominus} = -900$ kJ

Reaction 2 $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$ $\Delta H^{\ominus} = -91$ kJ

a) In industry these reactions are carried out in the presence of catalysts. A platinum catalyst is used in *Reaction 1* and copper catalyst is used in *Reaction 2*.

State and explain the effect on the yield of a reaction when a catalyst is used. (2 marks)

No effect. (1)

A catalyst increases the rates of both the forward reaction and the backward reaction to the same extent. (1)

b) State and explain which of the above reactions will give an increase in the yield of product(s) when the pressure is increased at a constant temperature. (3 marks)

Reaction 2 (1)

An increase in pressure will bring about a net reaction that decreases the number of moles of gas. This helps to reduce the pressure. (1)

The position of equilibrium will shift to the side of the equation with a fewer number of moles of gas, i.e. the product side of *Reaction 2*. (1)

Thus, the yield of the product in *Reaction 2* will increase.

c) State and explain the effect on the yield of NO(g) when the temperature is increased in *Reaction 1* at a constant pressure. (3 marks)

When the temperature is increased, the system will respond by reducing the temperature. (1)

As the backward reaction is endothermic, the system will undergo a net backward reaction. (1)

Thus, the yield of NO(g) will decrease. (1)

d) i) Write an expression for the equilibrium constant, K_c, for *Reaction 2*. State the units of K_c. (2 marks)

$$K_c = \frac{[CH_3OH(g)]}{[CO(g)][H_2(g)]^2}$$ (1)

Units of K_c : none (1)

5.

a) State Le Chatelier's principle. (2 marks)

Le Chatelier's principle states that if the conditions of a system in equilibrium is changed, the position of equilibrium will shift (1)

so as to reduce that change. (1)