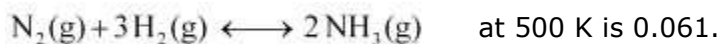


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(Class – XI)

## Question 7.21:

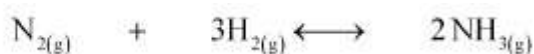
Equilibrium constant,  $K_c$  for the reaction



At a particular time, the analysis shows that composition of the reaction mixture is  $3.0 \text{ mol L}^{-1} \text{ N}_2$ ,  $2.0 \text{ mol L}^{-1} \text{ H}_2$  and  $0.5 \text{ mol L}^{-1} \text{ NH}_3$ . Is the reaction at equilibrium? If not in which direction does the reaction tend to proceed to reach equilibrium?

## Answer 7.21:

The given reaction is:



At a particular time:  $3.0 \text{ mol L}^{-1}$        $2.0 \text{ mol L}^{-1}$        $0.5 \text{ mol L}^{-1}$

Now, we know that,

$$\begin{aligned} Q_c &= \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} \\ &= \frac{(0.5)^2}{(3.0)(2.0)^3} \\ &= 0.0104 \end{aligned}$$

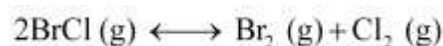
It is given that  $K_c = 0.061$ .

Since  $Q_c \neq K_c$  , the reaction is not at equilibrium.

Since  $Q_c < K_c$  , the reaction will proceed in the forward direction to reach equilibrium.

## Question 7.22:

Bromine monochloride, BrCl decomposes into bromine and chlorine and reaches the equilibrium:



for which  $K_c = 32$  at 500 K.

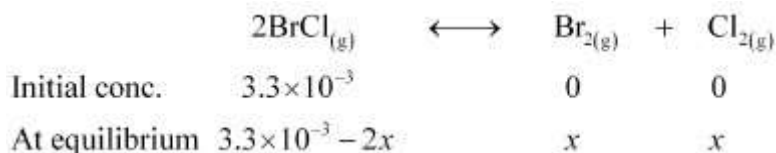
If initially pure BrCl is present at a concentration of  $3.3 \times 10^{-3} \text{ mol L}^{-1}$ , what is its molar concentration in the mixture at equilibrium?

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## Answer 7.22:

Let the amount of bromine and chlorine formed at equilibrium be  $x$ . The given reaction is:



Now, we can write,

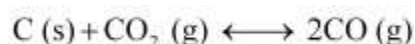
$$\begin{aligned} \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2} &= K_c \\ \Rightarrow \frac{x \times x}{(3.3 \times 10^{-3} - 2x)^2} &= 32 \\ \Rightarrow \frac{x}{3.3 \times 10^{-3} - 2x} &= 5.66 \\ \Rightarrow x &= 18.678 \times 10^{-3} - 11.32x \\ \Rightarrow 12.32x &= 18.678 \times 10^{-3} \\ \Rightarrow x &= 1.5 \times 10^{-3} \end{aligned}$$

Therefore, at equilibrium,

$$\begin{aligned} [\text{BrCl}] &= 3.3 \times 10^{-3} - (2 \times 1.5 \times 10^{-3}) \\ &= 3.3 \times 10^{-3} - 3.0 \times 10^{-3} \\ &= 0.3 \times 10^{-3} \\ &= 3.0 \times 10^{-4} \text{ molL}^{-1} \end{aligned}$$

## Question 7.23:

At 1127 K and 1 atm pressure, a gaseous mixture of CO and CO<sub>2</sub> in equilibrium with solid carbon has 90.55% CO by mass



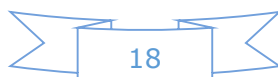
Calculate  $K_c$  for this reaction at the above temperature.

## Answer 7.23:

Let the total mass of the gaseous mixture be 100 g.

Mass of CO = 90.55 g

And, mass of CO<sub>2</sub> = (100 - 90.55) = 9.45 g



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Now, number of moles of CO,  $n_{\text{CO}} = \frac{90.55}{28} = 3.234 \text{ mol}$

Number of moles of CO<sub>2</sub>,  $n_{\text{CO}_2} = \frac{9.45}{44} = 0.215 \text{ mol}$

Partial pressure of CO,

$$\begin{aligned} p_{\text{CO}} &= \frac{n_{\text{CO}}}{n_{\text{CO}} + n_{\text{CO}_2}} \times p_{\text{total}} \\ &= \frac{3.234}{3.234 + 0.215} \times 1 \\ &= 0.938 \text{ atm} \end{aligned}$$

Partial pressure of CO<sub>2</sub>,

$$\begin{aligned} p_{\text{CO}_2} &= \frac{n_{\text{CO}_2}}{n_{\text{CO}} + n_{\text{CO}_2}} \times p_{\text{total}} \\ &= \frac{0.215}{3.234 + 0.215} \times 1 \\ &= 0.062 \text{ atm} \end{aligned}$$

$$\begin{aligned} \text{Therefore, } K_p &= \frac{[\text{CO}]^2}{[\text{CO}_2]} \\ &= \frac{(0.938)^2}{0.062} \\ &= 14.19 \end{aligned}$$

For the given reaction,

$$\Delta n = 2 - 1 = 1$$

We know that,

$$\begin{aligned} K_p &= K_c (RT)^{\Delta n} \\ \Rightarrow 14.19 &= K_c (0.082 \times 1127)^1 \\ \Rightarrow K_c &= \frac{14.19}{0.082 \times 1127} \\ &= 0.154 (\text{approximately}) \end{aligned}$$

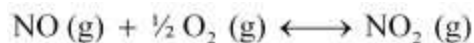
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## Question 7.24:

Calculate

- a)  $\Delta G^\circ$  and  
b) the equilibrium constant for the formation of  $\text{NO}_2$  from  $\text{NO}$  and  $\text{O}_2$  at 298 K



Where:

$$\Delta_f G^\circ (\text{NO}_2) = 52.0 \text{ kJ/mol}$$

$$\Delta_f G^\circ (\text{NO}) = 87.0 \text{ kJ/mol}$$

$$\Delta_f G^\circ (\text{O}_2) = 0 \text{ kJ/mol}$$

## Answer 7.24:

(a) For the given reaction,

$$\Delta G^\circ = \Delta G^\circ(\text{Products}) - \Delta G^\circ(\text{Reactants})$$

$$\Delta G^\circ = 52.0 - \{87.0 + 0\}$$

$$= -35.0 \text{ kJ mol}^{-1}$$

(b) We know that,

$$\Delta G^\circ = RT \log K_c$$

$$\Delta G^\circ = 2.303 RT \log K_c$$

$$K_c = \frac{-35.0 \times 10^{-3}}{-2.303 \times 8.314 \times 298}$$
$$= 6.134$$

$$\therefore K_c = \text{antilog} (6.134)$$
$$= 1.36 \times 10^6$$

Hence, the equilibrium constant for the given reaction  $K_c$  is  $1.36 \times 10^6$

## Question 7.25:

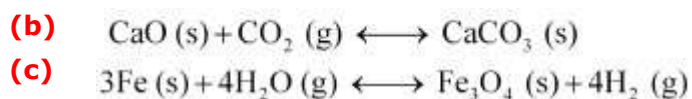
Does the number of moles of reaction products increase, decrease or remain same when each of the following equilibria is subjected to a decrease in pressure by increasing the volume?



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## Answer 7.25:

- (a) The number of moles of reaction products will increase. According to Le Chatelier's principle, if pressure is decreased, then the equilibrium shifts in the direction in which the number of moles of gases is more. In the given reaction, the number of moles of gaseous products is more than that of gaseous reactants. Thus, the reaction will proceed in the forward direction. As a result, the number of moles of reaction products will increase.
- (b) The number of moles of reaction products will decrease.
- (c) The number of moles of reaction products remains the same.

## Question 7.26:

Which of the following reactions will get affected by increasing the pressure?

Also, mention whether change will cause the reaction to go into forward or backward direction.

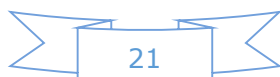
- (i)  $\text{COCl}_2 \text{ (g)} \rightleftharpoons \text{CO (g)} + \text{Cl}_2 \text{ (g)}$   
(ii)  $\text{CH}_4 \text{ (g)} + 2\text{S}_2 \text{ (g)} \rightleftharpoons \text{CS}_2 \text{ (g)} + 2\text{H}_2\text{S (g)}$   
(iii)  $\text{CO}_2 \text{ (g)} + \text{C (s)} \rightleftharpoons 2\text{CO (g)}$   
(iv)  $2\text{H}_2 \text{ (g)} + \text{CO (g)} \rightleftharpoons \text{CH}_3\text{OH (g)}$   
(v)  $\text{CaCO}_3 \text{ (s)} \rightleftharpoons \text{CaO (s)} + \text{CO}_2 \text{ (g)}$   
(vi)  $4\text{NH}_3 \text{ (g)} + 5\text{O}_2 \text{ (g)} \rightleftharpoons 4\text{NO (g)} + 6\text{H}_2\text{O (g)}$

## Answer 7.26:

The reactions given in (i), (iii), (iv), (v), and (vi) will get affected by increasing the pressure.

The reaction given in (iv) will proceed in the forward direction because the number of moles of gaseous reactants is more than that of gaseous products.

The reactions given in (i), (iii), (v), and (vi) will shift in the backward direction because the number of moles of gaseous reactants is less than that of gaseous products.

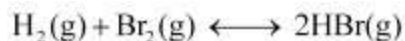


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## Question 7.27:

The equilibrium constant for the following reaction is  $1.6 \times 10^5$  at 1024 K.



Find the equilibrium pressure of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024 K.

## Answer 7.27:

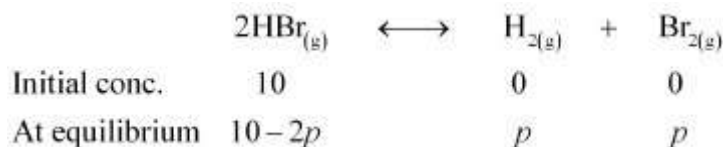
Given,

$K_p$  for the reaction i.e.,  $\text{H}_{2(\text{g})} + \text{Br}_{2(\text{g})} \rightleftharpoons 2\text{HBr}_{(\text{g})}$  is  $1.6 \times 10^5$ .

Therefore, for the reaction  $2\text{HBr}_{(\text{g})} \rightleftharpoons \text{H}_{2(\text{g})} + \text{Br}_{2(\text{g})}$ , the equilibrium constant will be,

$$\begin{aligned} K'_p &= \frac{1}{K_p} \\ &= \frac{1}{1.6 \times 10^5} \\ &= 6.25 \times 10^{-6} \end{aligned}$$

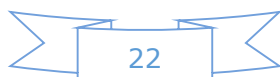
Now, let  $p$  be the pressure of both  $\text{H}_2$  and  $\text{Br}_2$  at equilibrium.



Now, we can write,

$$\begin{aligned} \frac{P_{\text{HBr}} \times P_{\text{Br}_2}}{P_{\text{HBr}}^2} &= K'_p \\ \frac{p \times p}{(10 - 2p)^2} &= 6.25 \times 10^{-6} \\ \frac{p}{10 - 2p} &= 2.5 \times 10^{-3} \\ p &= 2.5 \times 10^{-2} - (5.0 \times 10^{-3})p \\ p + (5.0 \times 10^{-3})p &= 2.5 \times 10^{-2} \\ (1005 \times 10^{-3})p &= 2.5 \times 10^{-2} \\ p &= 2.49 \times 10^{-2} \text{ bar} = 2.5 \times 10^{-2} \text{ bar (approximately)} \end{aligned}$$

Therefore, at equilibrium,



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$$\begin{aligned}[\text{H}_2] &= [\text{Br}_2] = 2.49 \times 10^{-2} \text{ bar} \\ [\text{HBr}] &= 10 - 2 \times (2.49 \times 10^{-2}) \text{ bar} \\ &= 9.95 \text{ bar} = 10 \text{ bar (approximately)}\end{aligned}$$

## Question 7.28:

Dihydrogen gas is obtained from natural gas by partial oxidation with steam as per following endothermic reaction:



- (a) Write an expression for  $K_p$  for the above reaction.
- (b) How will the values of  $K_p$  and composition of equilibrium mixture be affected by
- Increasing the pressure
  - Increasing the temperature
  - Using a catalyst?

## Answer 7.28:

(a) For the given reaction,

$$K_p = \frac{P_{\text{CO}} \times P_{\text{H}_2}^3}{P_{\text{CH}_4} \times P_{\text{H}_2\text{O}}}$$

- (b)
- According to Le Chatelier's principle, the equilibrium will shift in the backward direction.
  - According to Le Chatelier's principle, as the reaction is endothermic, the equilibrium will shift in the forward direction.
  - The equilibrium of the reaction is not affected by the presence of a catalyst. A catalyst only increases the rate of a reaction. Thus, equilibrium will be attained quickly.

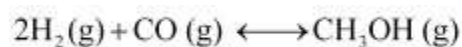
## Question 7.29:

Describe the effect of:

- Addition of  $\text{H}_2$
- Addition of  $\text{CH}_3\text{OH}$
- Removal of  $\text{CO}$
- Removal of  $\text{CH}_3\text{OH}$  on the equilibrium of the reaction:

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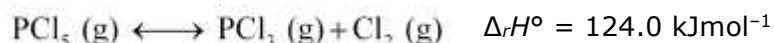


## Answer 7.29:

- (a) According to Le Chatelier's principle, on addition of  $\text{H}_2$ , the equilibrium of the given reaction will shift in the forward direction.
- (b) On addition of  $\text{CH}_3\text{OH}$ , the equilibrium will shift in the backward direction.
- (c) On removing  $\text{CO}$ , the equilibrium will shift in the backward direction.
- (d) On removing  $\text{CH}_3\text{OH}$ , the equilibrium will shift in the forward direction.

## Question 7.30:

At 473 K, equilibrium constant  $K_c$  for decomposition of phosphorus pentachloride,  $\text{PCl}_5$  is  $8.3 \times 10^{-3}$ . If decomposition is depicted as,



- a) Write an expression for  $K_c$  for the reaction.
- b) What is the value of  $K_c$  for the reverse reaction at the same temperature?
- c) What would be the effect on  $K_c$  if
  - (i) more  $\text{PCl}_5$  is added
  - (ii) pressure is increased?
  - (iii) The temperature is increased?

## Answer 7.30:

(a) 
$$K_c = \frac{[\text{PCl}_3(\text{g})][\text{Cl}_2(\text{g})]}{[\text{PCl}_5(\text{g})]}$$

- (b) Value of  $K_c$  for the reverse reaction at the same temperature is:

$$\begin{aligned} K_c' &= \frac{1}{K_c} \\ &= \frac{1}{8.3 \times 10^{-3}} = 1.2048 \times 10^2 \\ &= 120 - 48 \end{aligned}$$

- (c) (i)  $K_c$  would remain the same because in this case, the temperature remains the same.  
(ii)  $K_c$  is constant at constant temperature. Thus, in this case,  $K_c$  would not change.



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(iii) In an endothermic reaction, the value of  $K_c$  increases with an increase in temperature.

Since the given reaction is an endothermic reaction, the value of  $K_c$  will increase if the temperature is increased.

## Question 7.31:

Dihydrogen gas used in Haber's process is produced by reacting methane from natural gas with high temperature steam. The first stage of two stage reaction involves the formation of CO and  $H_2$ . In second stage, CO formed in first stage is reacted with more steam in water gas shift reaction,

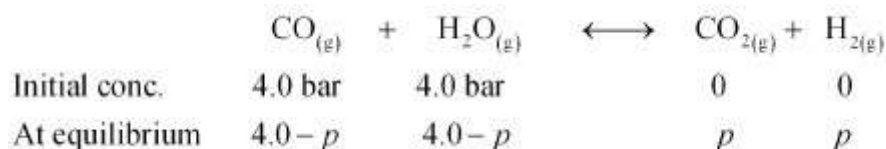


If a reaction vessel at  $400^\circ C$  is charged with an equimolar mixture of CO and steam such that  $P_{CO} = P_{H_2O} = 4.0$  bar, what will be the partial pressure of  $H_2$  at equilibrium?

$K_p = 10.1$  at  $400^\circ C$

## Answer 7.31:

Let the partial pressure of both carbon dioxide and hydrogen gas be  $p$ . The given reaction is:



It is given that  $K_p = 10.1$ .

Now,

$$\begin{aligned} \frac{P_{CO_2} \times P_{H_2}}{P_{CO} \times P_{H_2O}} &= K_p \\ \Rightarrow \frac{p \times p}{(4.0 - p)(4.0 - p)} &= 10.1 \\ \Rightarrow \frac{p}{4.0 - p} &= 3.178 \\ \Rightarrow p &= 12.712 - 3.178p \\ \Rightarrow 4.178p &= 12.712 \\ \Rightarrow p &= 3.04 \end{aligned}$$

Hence, at equilibrium, the partial pressure of  $H_2$  will be 3.04 bar.

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## Question 7.32:

Predict which of the following reaction will have appreciable concentration of reactants and products:

- a)  $\text{Cl}_2(\text{g}) \rightleftharpoons 2\text{Cl}(\text{g}); K_c = 5 \times 10^{-39}$   
b)  $\text{Cl}_2(\text{g}) + 2\text{NO}(\text{g}) \rightleftharpoons 2\text{NOCl}(\text{g}); K_c = 3.7 \times 10^8$   
c)  $\text{Cl}_2(\text{g}) + 2\text{NO}_2(\text{g}) \rightleftharpoons 2\text{NO}_2\text{Cl}(\text{g}); K_c = 1.8$

## Answer 7.32:

If the value of  $K_c$  lies between  $10^{-3}$  and  $10^3$ , a reaction has appreciable concentration of reactants and products. Thus, the reaction given in (c) will have appreciable concentration of reactants and products.

## Question 7.33:

The value of  $K_c$  for the reaction  $3\text{O}_2(\text{g}) \rightleftharpoons 2\text{O}_3(\text{g})$  is  $2.0 \times 10^{-50}$  at  $25^\circ\text{C}$ . If the equilibrium concentration of  $\text{O}_2$  in air at  $25^\circ\text{C}$  is  $1.6 \times 10^{-2}$ , what is the concentration of  $\text{O}_3$ ?

## Answer 7.33:

The given reaction is:



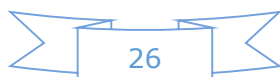
$$\text{Then, } K_c = \frac{[\text{O}_{3(\text{g})}]^2}{[\text{O}_{2(\text{g})}]^3}$$

$$\text{It is given that } K_c = 2.0 \times 10^{-50} \text{ and } [\text{O}_{2(\text{g})}] = 1.6 \times 10^{-2}.$$

Then, we have,

$$\begin{aligned} 2.0 \times 10^{-50} &= \frac{[\text{O}_{3(\text{g})}]^2}{[1.6 \times 10^{-2}]^3} \\ \Rightarrow [\text{O}_{3(\text{g})}]^2 &= 2.0 \times 10^{-50} \times (1.6 \times 10^{-2})^3 \\ \Rightarrow [\text{O}_{3(\text{g})}]^2 &= 8.192 \times 10^{-56} \\ \Rightarrow [\text{O}_{3(\text{g})}] &= 2.86 \times 10^{-28} \text{ M} \end{aligned}$$

Hence, the concentration of  $\text{O}_3$  is  $2.86 \times 10^{-28}$  M.



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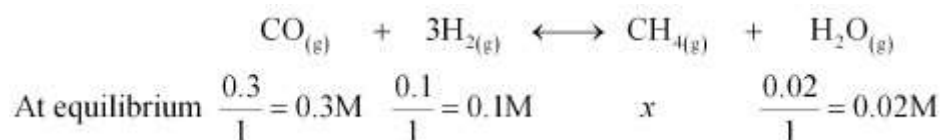
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## Question 7.34:

The reaction,  $\text{CO}(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g})$  is at equilibrium at 1300 K in a 1L flask. It also contain 0.30 mol of CO, 0.10 mol of  $\text{H}_2$  and 0.02 mol of  $\text{H}_2\text{O}$  and an unknown amount of  $\text{CH}_4$  in the flask. Determine the concentration of  $\text{CH}_4$  in the mixture. The equilibrium constant,  $K_c$  for the reaction at the given temperature is 3.90.

## Answer 7.34:

Let the concentration of methane at equilibrium be  $x$ .



It is given that  $K_c = 3.90$ .

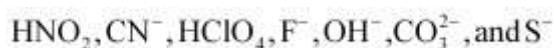
Therefore,

$$\begin{aligned} \frac{[\text{CH}_{4(\text{g})}][\text{H}_2\text{O}_{(\text{g})}]}{[\text{CO}_{(\text{g})}][\text{H}_{2(\text{g})}]^3} &= K_c \\ \Rightarrow \frac{x \times 0.02}{0.3 \times (0.1)^3} &= 3.90 \\ \Rightarrow x &= \frac{3.90 \times 0.3 \times (0.1)^3}{0.02} \\ &= \frac{0.00117}{0.02} \\ &= 0.0585\text{M} \\ &= 5.85 \times 10^{-2}\text{M} \end{aligned}$$

Hence, the concentration of  $\text{CH}_4$  at equilibrium is  $5.85 \times 10^{-2}\text{M}$ .

## Question 7.35:

What is meant by the conjugate acid-base pair? Find the conjugate acid/base for the following species:



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## Answer 7.35:

A conjugate acid-base pair is a pair that differs only by one proton. The conjugate acid-base for the given species is mentioned in the table below.

### Species Conjugate acid-base

$\text{HNO}_2$	$\text{NO}_2^-$ (base)
$\text{CN}^-$	$\text{HCN}$ (acid)
$\text{HClO}_4$	$\text{ClO}_4^-$ (base)
$\text{F}^-$	$\text{HF}$ (acid)
$\text{OH}^-$	$\text{H}_2\text{O}$ (acid) / $\text{O}^{2-}$ (base)
$\text{CO}_3^{2-}$	$\text{HCO}_3^-$ (acid)
$\text{S}^{2-}$	$\text{HS}^-$ (acid)

## Question 7.36:

Which of the followings are Lewis acids?  $\text{H}_2\text{O}$ ,  $\text{BF}_3$ ,  $\text{H}^+$ , and  $\text{NH}_4^+$

## Answer 7.36:

Lewis acids are those acids which can accept a pair of electrons. For example,  $\text{BF}_3$ ,  $\text{H}^+$ , and  $\text{NH}_4^+$  are Lewis acids.

## Question 7.37:

What will be the conjugate bases for the Brønsted acids:  $\text{HF}$ ,  $\text{H}_2\text{SO}_4$  and  $\text{HCO}_3^-$ ?

## Answer 7.37:

The table below lists the conjugate bases for the given Brønsted acids.

### Bronsted acid Conjugate base

$\text{HF}$	$\text{F}^-$
$\text{H}_2\text{SO}_4$	$\text{HSO}_4^-$
$\text{HCO}_3^-$	$\text{CO}_3^{2-}$

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## Question 7.38:

Write the conjugate acids for the following Brønsted bases:  $\text{NH}_2^-$ ,  $\text{NH}_3$  and  $\text{HCOO}^-$ .

## Answer 7.38:

The table below lists the conjugate acids for the given Bronsted bases.

### Bronsted base Conjugate acid

$\text{NH}_2^-$	$\text{NH}_3$
$\text{NH}_3$	$\text{NH}_4^+$
$\text{HCOO}^-$	$\text{HCOOH}$

## Question 7.39:

The species:  $\text{H}_2\text{O}$ ,  $\text{HCO}_3^-$ ,  $\text{HSO}_4^-$ , and  $\text{NH}_3$  can act both as Brønsted acids and bases. For each case give the corresponding conjugate acid and base.

## Answer 7.39:

The table below lists the conjugate acids and conjugate bases for the given species.

### Species Conjugate acid Conjugate base

$\text{H}_2\text{O}$	$\text{H}_3\text{O}^+$	$\text{OH}^-$
$\text{HCO}_3^-$	$\text{H}_2\text{CO}_3$	$\text{CO}_3^{2-}$
$\text{HSO}_4^-$	$\text{H}_2\text{SO}_4$	$\text{SO}_4^{2-}$
$\text{NH}_3$	$\text{NH}_4^+$	$\text{NH}_2^-$

# Chemistry

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(Chapter – 7) (Equilibrium)  
(Class – XI)

## Question 7.40:

Classify the following species into Lewis acids and Lewis bases and show how these act as Lewis acid/base:

- (a)  $\text{OH}^-$
- (b)  $\text{F}^-$
- (c)  $\text{H}^+$
- (d)  $\text{BCl}_3$ .

## Answer 7.40:

- (a)  $\text{OH}^-$  is a Lewis base since it can donate its lone pair of electrons.
- (b)  $\text{F}^-$  is a Lewis base since it can donate a pair of electrons.
- (c)  $\text{H}^+$  is a Lewis acid since it can accept a pair of electrons.
- (d)  $\text{BCl}_3$  is a Lewis acid since it can accept a pair of electrons.

