

## Gibbs Free Energy Chemistry Questions with Solutions

**Q1. If the reaction quotient (Q) is greater than the equilibrium constant (K), what is true about the Gibbs free energy?**

- a.) It is greater than zero.
- b.) It is less than zero.
- c.) More information is needed to determine the Gibbs free energy.
- d.) It is equal to zero.

**Correct Answer–** (a.) It is greater than zero.

**Explanation–** If Q is greater than K, the reaction has exceeded the equilibrium state. It will proceed non-spontaneously, and this means that the  $\Delta G$  must be positive, or greater than zero.

**Q2. The standard reaction Gibbs energy for a chemical reaction at an absolute temperature T is given by:  $\Delta_r G^\circ = A - BT$**

**Where A and B are non-zero constants. Choose the correct statement about this reaction?**

- a.) The reaction is exothermic if  $B < 0$ .
- b.) The reaction is exothermic if  $A > 0$  and  $B < 0$ .
- c.) The reaction is endothermic if  $A < 0$  and  $B > 0$ .
- d.) The reaction is endothermic if  $A > 0$ .

**Correct Answer–** (d.) The reaction is endothermic if  $A > 0$ .

**Explanation–**  $\Delta G = \Delta H - T\Delta S$

$\Delta H = A$

When  $\Delta H = +ve$ , the reaction will be endothermic.

**Q3. Calculate the Gibbs free energy for the reaction of conversion of ATP into ADP at 293 Kelvin the change in enthalpy is 19.07 Kcal and the change in entropy is 90 cal per Kelvin.**

- a.) 7.3 Cal
- b.) -5.3Kcal
- c.) 7.3 Kcal
- d.) -7.3Kcal

**Correct Answer–** (d.) -7.3Kcal.

**Explanation–**  $\Delta G = \Delta H - T\Delta S$ ; here  $\Delta H = 19.07$  kcal and  $\Delta S = 90$  cal/K,

$\Delta G = 19.07$  Kcal -  $293(90$  cal/K)

=  $19.07$  Kcal -  $26.37$  Kcal

=  $-7300$  cal =  $-7.3$  Kcal.

The Gibbs free energy change is  $-7.3$  Kcal.

**Q4. Gibbs energy change  $\Delta G$  is related to equilibrium constant K as:**

- a.)  $\Delta G^\circ = -RT \ln k$
- b.)  $\Delta G^\circ = RT \ln k$
- c.)  $\ln k = -RT/\Delta G^\circ$
- d.)  $\ln k = \Delta G^\circ/RT$

**Correct Answer–** (a.)  $\Delta G^\circ = -RT \ln k$

**Q5. Which is not the correct relationship between  $\Delta G^\circ$  and equilibrium constant  $K_p$ ?**

- a.)  $K_p = -RT \log \Delta G^\circ$
- b.)  $K_p = [e/RT] \Delta G^\circ$
- c.)  $K_p = -\Delta G^\circ/RT$
- d.)  $K_p = e^{-\Delta G^\circ/RT}$

**Correct Answer–** (a.)  $K_p = -RT \log \Delta G^\circ$ , (b.)  $K_p = [e/RT] \Delta G^\circ$ , (c.)  $K_p = -\Delta G^\circ/RT$

**Q6. Fill in the blank.**

**The Gibbs free energy is positive when a change in enthalpy and change in entropy is positive at \_\_\_\_.**

**Answer.** The Gibbs free energy is positive when a change in enthalpy and change in entropy is positive at low temperatures.  
It is negative at high temperatures.

**Q7. Is Gibbs free energy always negative?**

**Answer.** Gibbs free energy is not always negative. When the reaction is spontaneous at a given temperature, it is negative, and for the non-spontaneous reaction, Gibbs free energy is positive.

**Q8. State True or False.**

**$\Delta G = \Delta G^\circ + RT \log K$ .**

**Answer.** True.

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = \Delta G^\circ + RT \log K$$

At equilibrium,  $\Delta G = 0$

Therefore,  $\Delta G^\circ = -RT \log K$ .

**Q9. When is the value of the equilibrium constant less than 1?**

**Answer.** The value of the equilibrium constant is less than 1 when the reaction is non-spontaneous. It means that the equilibrium mixture contains more reactants than products. Therefore the value of  $\Delta G^\circ > 0$ .

**Q10. What is Gibbs free energy for a reversible reaction at equilibrium?**

**Answer.** Gibbs free energy is given by  $\Delta G = \Delta H - T\Delta S$   
For a reaction to be reversible,  $\Delta G$  must be equal to 0.

**Q11. Give the correct relation between equilibrium constant (K), standard free energy ( $\Delta G^\circ$ ) and temperature (T).**

**Answer.** The correct relation between equilibrium constant (K), standard free energy ( $\Delta G^\circ$ ) and temperature (T) can be given as-

Consider a reaction,  $A + B \rightleftharpoons C + D$

$$\Delta G = \Delta G^\circ + RT \ln K$$

For equilibrium,  $\Delta G = 0$

$$0 = \Delta G^\circ + RT \ln K$$

$$\Delta G^\circ = -RT \ln K$$

$$\Delta G^\circ = -2.303RT \log K$$

$$K = 10^{-\Delta G^\circ / 2.303RT}$$

**Q12. Calculate the Standard Free Energy Change at 25°C given the Equilibrium constant of  $1.3 \times 10^4$ .**

**Answer.**  $\Delta G^\circ = -2.303 RT \log K$

$$\Delta G^\circ = -2.303 RT \log K = -8.314 \times 298 \times \log 1.3 \times 10^4 = -23469 \text{ J} = -23.4 \text{ KJ}$$

**Q13. Using the Gibbs change,  $\Delta G^\circ = +63.3 \text{ kJ}$  for the reaction  $\text{Ag}_2\text{CO}_3 \rightleftharpoons 2\text{Ag}^+ + \text{CO}_3^{2-}$ . Calculate the  $K_{\text{sp}}$  for  $\text{Ag}_2\text{CO}_3$  in water at 25°C.**

**Answer.** The relationship between solubility product and Gibbs free energy is-

$$\Delta G^\circ = -2.303 RT \log K_{\text{sp}}$$

On substituting the values-

$$63.3 \times 10^3 = -2.303 \times 8.31 \times 298 \log K_{\text{sp}}$$

$$-11.09 = \log K_{\text{sp}}$$

$$8 \times 10^{-12} = K_{\text{sp}}$$

Hence, the  $K_{\text{sp}}$  for  $\text{Ag}_2\text{CO}_3$  in water at 25°C is  $8 \times 10^{-12}$ .

**Q14. Calculate the standard free energy at 1 atm 60°C for  $\text{N}_2\text{O}_4$ . It is 50% dissociated at this temperature.**

**Answer.**  $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$

If 50% is dissociated, then the mole fraction of both substances will be-

$$X_{N_2O_4} = \frac{1.05}{1 + 1.05}$$

$$X_{NO_2} = \frac{2 \times .05}{1 + 1.05}$$

Pressure-

$$P_{N_2O_4} = \frac{0.5}{1.5} \times 1 \text{ atm}$$

$$P_{NO_2} = \frac{1}{1.5} \times 1 \text{ atm}$$

$$K_P = P_{NO_2}/P_{N_2O_4} = 1.5/(1.5)^2(0.5) = 1.33$$

$$\Delta G^\circ = -2.303 RT \log K_P,$$

$$T=333 \text{ K}$$

$$\Delta G^\circ = -2.303 \times 8.314 \times 333 \log (1.33)$$

$$\Delta G^\circ = -789.34 \text{ KJ/mol.}$$

**Q15. Is the reaction rate affected by Gibbs free energy?**

**Answer.** Concentration, pressure, temperature, and surface area are all factors that influence reaction rate.

There are two variables in the Gibbs Free Energy equation-

$\Delta H$ , which is a system's enthalpy, and  $\Delta S$ , which is its entropy.

Gibbs Free Energy indicates whether a chemical change is thermodynamically possible.

The Gibbs Free energy  $\Delta G$  must be -ve for a reaction to be spontaneous.  $\Delta H < 0$  and  $\Delta S > 0$  to attain this.

As a result, the temperature and entropy that affect Gibbs Free Energy also affect the reaction rate indirectly.

Hence, we can say that the reaction rate is affected by Gibbs energy.

## Practise Questions on Gibbs Free Energy

**Q1. If the reaction quotient (Q) is less than the equilibrium constant (K). Choose the correct statement about Gibbs free energy?**

- a.) We must know the reaction enthalpy to determine it.
- b.) It is greater than zero.
- c.) It is equal to zero.
- d.) It is less than zero

**Correct Answer-** (d.) It is less than zero.

**Explanation**– If  $Q$  is less than  $K$ , then the reaction has not yet reached the equilibrium state. It will proceed spontaneously in the forward direction, and this must mean that the  $\Delta G$  must be negative, or less than zero.

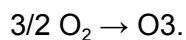
**Q2. What is the relation between Gibbs free energy and the EMF of the cell?**

- a.)  $\Delta G = -nFE_{\text{cell}}$
- b.)  $G = -nFE_{\text{cell}}$
- c.)  $\Delta G = -nE_{\text{cell}}$
- d.)  $\Delta G = -nF$

**Correct Answer**– (a.)  $\Delta G = -nFE_{\text{cell}}$

**Q3. Calculate the Gibbs free energy for the conversion of oxygen to Ozone at room temperature if  $K_p$  is given as  $2.47 \times 10^{-29}$ .**

**Answer.** The chemical equation for the conversion of oxygen to Ozone is

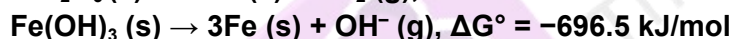
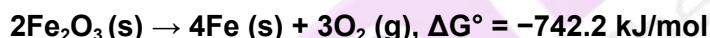


$$\Delta G = -2.303 RT \log K_p$$

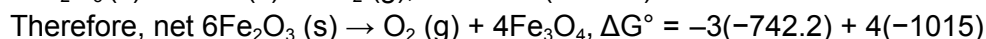
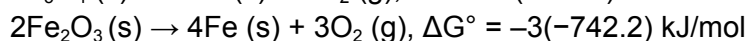
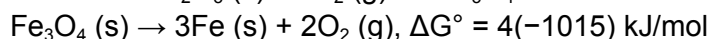
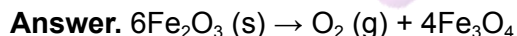
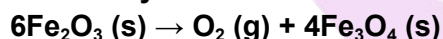
By substituting,

$$\Delta G = -2.303 \times 8.314 \text{ J/K-mol} \times 293\text{K} \times 2.47 \times 10^{-29} = 163000 \text{ J/mol} = 16.3 \text{ kJ/mol.}$$

**Q4. Use the given the standard Gibbs energy changes for these equations:**



**To identify the  $\Delta G^\circ$  for the following reaction**



$$\Delta G^\circ = 3837.4 \text{ KJ/mol.}$$

**Q5. Calculate  $\Delta G^\circ$  for the reaction  $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$  at 298.15 K.**

**Given  $\Delta H = -128.3 \text{ kJ}$ ,  $\Delta S = -159.5 \text{ JK}^{-1}$**

**Answer.**  $\Delta G^\circ = \Delta H - T\Delta S$

$$\Delta G^\circ = -128.3 - \{(298.15) \times -[159.5 \text{ JK}^{-1} \times (1\text{J}/1000\text{kJ})]\}$$

$$\Delta G^\circ = -80.75 \text{ kJ.}$$