Marks
9

- In an experiment, NOCl (2.00 mol) was placed in a closed 1.00 L flask. After equilibrium was established at 25 °C, the concentration of NO(g) was 0.66 M. Calculate the value of K_c at 25 °C for the following reaction.

$$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$$

K_c =

Calculate the value of K_p at 25 °C for the reaction above.

K_p =

Given that $\Delta H_f°$ for NOCl(g) = 51.71 kJ mol^{-1} and $\Delta H_f°$ for NO(g) = 90.29 kJ mol^{-1} at 25 °C, calculate the value of $\Delta H°$ for the reaction above.

$\Delta H°_{rxn}$ =

What is the effect upon the [NOCl] of an equilibrium mixture if the temperature is increased?

In which direction will the equilibrium shift if the volume of the flask is reduced?

Marks
5

- Equal volumes of carbon monoxide and hydrogen gas are introduced into a sealed 4.5 L flask at 1200 K and the following equilibrium is established.

$$CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g) \qquad \Delta H° = -205.9 \text{ kJ mol}^{-1}$$

At equilibrium, the flask contains 0.22 mol of CH_4 and the total pressure in the flask is 46.4 atm. Calculate the amount of $H_2(g)$ (in mol) that was initially introduced into the flask.

Answer:

In a separate experiment, it is determined that the reaction is in equilibrium when the same 4.5 L flask contains 0.18 mol of CH_4, 0.24 mol of H_2O, 0.82 mol of CO and 0.65 mol of H_2 at 1200 K. Calculate the concentration equilibrium constant, K_c, for this temperature.

K_c =

Marks 7

Calculate the partial pressure equilibrium constant, K_p, at 1200 K.

K_p =

What is the standard free energy change $\Delta G°$ for the forward reaction (in kJ mol^{-1}) at 1200 K?

$\Delta G°$=

What will be the effect on the equilibrium if CO(g) is injected into the flask, which maintains a constant volume.

What will be the effect on the equilibrium if the temperature is decreased?

What will be the effect on the equilibrium if the volume of the flask is decreased?

What will be the effect on the equilibrium if the walls of the flask are refrigerated so that liquid water condenses out?

Marks
6

- Consider the following equilibrium in the gas-phase at 35 ºC.

$$Cl_2(g) + 2NO(g) \rightleftharpoons 2NOCl(g) \qquad K_c = 6.25 \times 10^4\ M^{-1}$$

Equimolar amounts of NOCl(g) and Cl_2(g) are introduced into a sealed 1.00 L flask. When the system reaches equilibrium at 35 ºC, the concentration of NO(g) in the flask is 4.04×10^{-4} M. What amount of Cl_2(g) (in mol) was initially added to the flask?

Answer:

At the same temperature (35 ºC) O_2(g) reacts with NO(g) according to the equation:

$$O_2(g) + 2NO(g) \rightleftharpoons 2NO_2(g) \qquad K_c = 6.25\ M^{-1}$$

Determine K_c for the following reaction.

$$2NO_2(g) + Cl_2(g) \rightleftharpoons 2NOCl(g) + O_2(g)$$

K_c =

Marks
7

Calculate the partial pressure equilibrium constant, K_p, at 35 ºC for the reaction:

$$2NO_2(g) + Cl_2(g) \rightleftharpoons 2NOCl(g) + O_2(g)$$

K_p =

What is the standard free energy change, $\Delta G°$, for the forward reaction (in kJ mol^{-1}) at 35 ºC?

$\Delta G°$ =

If 0.150 mol of O_2(g) and 3.00 × 10^{-4} mol of NO_2(g) are added to the 1.00 L flask, determine the free energy change, ΔG, (in kJ mol^{-1}) as the system moves to its new equilibrium point.

ΔG =

Will the amount of NO_2(g) in the flask increase or decrease as the system moves to its new equilibrium position? Explain.

Marks
4

The CO(g) in water gas can be reacted further with H_2O(g) in the so-called "water-gas shift" reaction:

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

At 900 K, K_c = 1.56 for this reaction. A sample of water gas flowing over coal at 900 K contains a 1:1 mole ratio of CO(g) and H_2(g), as well as 0.250 mol L^{-1} H_2O(g). This sample is placed in a sealed container at 900 K and allowed to come to equilibrium, at which point it contains 0.070 mol L^{-1} CO_2(g). What was the initial concentration of CO(g) and H_2(g) in the sample?

[CO] = [H_2] =

If the walls of the container are chilled to below 100 ºC, what will be the effect on the concentration of CO_2(g)?

Marks
6

- The isomerisation of glucose-6-phosphate (G6P) to fructose-6-phosphate (F6P) is a key step in the metabolism of glucose for energy. At 298 K,

$$\text{G6P} \rightleftharpoons \text{F6P} \qquad \Delta G^\circ = 1.67 \text{ kJ mol}^{-1}$$

Calculate the equilibrium constant for this process at 298 K.

Answer:

What is the free energy change (in kJ mol^{-1}) involved in a mixture of 3.00 mol of F6P and 2.00 mol of G6P reaching equilibrium at 298 K?

Answer:

Sketch a graph of G_{sys} *versus* “extent of reaction”, with a curve showing how G_{sys} varies as G6P is converted to F6P. Indicate the position on this curve corresponding to 3.00 mol of F6P and 2.00 mol of G6P.

Marks
4

- Acetylene, C_2H_2, can be produced by reacting calcium carbide, CaC_2, with water:

$$CaC_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$$

A 1.000 g sample of CaC_2 is placed in a sealed vessel that contains 250.0 mL of H_2O(l) and 250.0 mL of N_2(g) at 1.000 atm, and allowed to react completely with the water. The final pressure in the sealed vessel at 22.0 ºC is 2.537 atm. Determine the vapour pressure of water in the sealed vessel at 22.0 ºC. Give your answer in mmHg. Ignore any change in the volume of the water.

Answer:

The solubility of acetylene in water at 22.0 ºC is small. If the temperature were raised, would you expect this solubility to increase or decrease?

Marks
5

- Consider the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

  $\Delta H^\circ = -198.4$ kJ mol^{-1} and $\Delta S^\circ = -187.9$ J K^{-1} mol^{-1} at 25 °C.

  Show that this reaction is spontaneous at 25 °C.

  If the volume of the reaction system is increased at 25 °C, in which direction will the reaction move?

  Calculate the value of the equilibrium constant, K, at 25 °C.

  K =

  Assuming ΔH° and ΔS° are independent of temperature, in which temperature range is the reaction non-spontaneous?

  Answer:

Marks
6

- The first step in the metabolism of glucose in biological systems is the addition of a phosphate group in a dehydration-condensation reaction:

$$\text{glucose(aq)} + H_2PO_4^-\text{(aq)} \rightleftharpoons \text{[glucose phosphate]}^-\text{(aq)} + H_2O\text{(l)}$$

The free energy change associated with this reaction is $\Delta G° = 13.8$ kJ mol^{-1}. The reaction is driven forwards by harnessing the free energy associated with the hydrolysis of adenosine triphosphate, ATP^{4-}, to adenosine diphosphate, ADP^{3-} :

$$ATP^{4-}\text{(aq)} + H_2O\text{(l)} \rightleftharpoons ADP^{3-}\text{(aq)} + H_2PO_4^-\text{(aq)} \qquad \Delta G° = -30.5 \text{ kJ mol}^{-1}$$

The overall reaction is thus:

$$\text{glucose(aq)} + ATP^{4-}\text{(aq)} \rightleftharpoons \text{[glucose phosphate]}^-\text{(aq)} + ADP^{3-}\text{(aq)}$$

Calculate the equilibrium constant associated with this overall reaction at body temperature (37 ºC).

Answer:

This overall equilibrium reaction is investigated by adding 0.0100 mol of ATP^{4-} to a flask containing 175 mL of a 0.0500 M aqueous solution of glucose at 37 ºC. What percentage of the ATP^{4-} will have been consumed when the system reaches equilibrium?

Answer: **87%**

Suggest two simple ways of further reducing the remaining percentage of ATP^{4-}.

Marks 5

- At 700 °C, hydrogen and iodine react according to the following equation.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad K_c = 49.0$$

Hydrogen also reacts with sulfur at 700 °C:

$$2H_2(g) + S_2(g) \rightleftharpoons 2H_2S(g) \qquad K_c = 1.075 \times 10^8$$

Determine K_c for the following overall equilibrium reaction at 700 °C.

$$2I_2(g) + 2H_2S(g) \rightleftharpoons S_2(g) + 4HI(g)$$

K_c =

What is the standard free energy change at 700 °C for this overall equilibrium reaction?

Answer:

Marks
5

If 0.250 mol of HI(g) is introduced into a 2.00 L flask at 700 °C, what will be the concentration of I_2(g) at equilibrium?

Answer:

If 0.274 g of H_2S were now introduced into the same flask, what would be the concentration of S_2(g) at equilibrium?

Answer:

	Marks
• Explain the meanings of the following terms.	**8**
Heat	
$P\Delta V$ work	
Internal energy	
Enthalpy change	
Entropy	
Equilibrium constant	
Reaction quotient	
Triple point	

- Consider the following reaction.

$$H_2O(g) + Cl_2O(g) \rightleftharpoons 2HOCl(g) \qquad K_p = 0.090 \text{ at } 298 \text{ K}$$

Calculate $\Delta G°$ (in J mol^{-1}) for this reaction.

3

$\Delta G°$ =

Calculate the reaction quotient, Q, at 25 °C when $p(H_2O)$ = 18 mmHg, $p(Cl_2O)$ = 2.0 mmHg and $p(HOCl)$ = 0.10 mmHg.

Q =

In which direction will the reaction proceed spontaneously at these partial pressures?