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Q1. Choose the correct answer.
A thermodynamic state function is a uantity
(i) used to determine heat changes
(ii) whose value is independent of path
(iii) used to determine pressure volume work
(iv) whose value depends on temperature only.

Q2. For the process to occur under adiabatic conditions, the correct condition is:
(i) $\Delta T = 0$ (ii) $\Delta p = 0$ (iii) $\Delta q = 0$ (iv) $w = 0$

Q3. The enthalpies of all elements in their standard states are:
(i) unity (ii) zero (iii) < 0 (iv) different for each element

Q4. $\Delta U^{\ominus}$ of combustion of methane is $-X$ kJ mol^{-1}. The value of $\Delta H^{\$}$ is:
(i) $= \Delta U^{\ominus}$ (ii) $> \Delta U^{\ominus}$ (iii) $< \Delta U^{\ominus}$ (iv) 0

Q5. A reaction, $A + B \longrightarrow C + D + q$ is found to have a positive entropy change. The reaction will be
(i) possible at high temperature
(ii) possible at low temperature
(iii) not possible at any temperature
(iv) possible at any temperature

Q6. For the reaction; $2Cl(g) \longrightarrow Cl_2(g)$; what will be the signs of ΔH and ΔS?

Q7. Comment on the thermodynamic stability of $NO(g)$ and $NO_2(g)$ given:

$$\frac{1}{2} N_2(g) + \frac{1}{2} O_2(g) \longrightarrow NO(g); \quad \Delta_f H^{\ominus} = 90 \text{ kJ mol}^{-1}$$

$$NO(g) + \frac{1}{2} O_2(g) \longrightarrow NO_2(g); \quad \Delta_f H^{\ominus} = -74 \text{ kJ mol}^{-1}$$

Q8. What happens to work, when
(a) gas expands against an external pressure.
(b) gas is compressed.
(c) gas expands into vacuum.
(d) an ideal gas expands reversibly and isothermally.

Q9. What are instant cold and hot packs? for what purpose they are used?

Q10. Give appropriate reasons
(a) It is preferable to determine $\Delta_r H$ than $\Delta_r U$
(b) Defining the standard state is necessary
(c) It is necessary to specify phases/ physical states of reactants and products while writing thermochemical equations.

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S1. As per the definition. The correct answer is (ii).

S2. The correct answer is (iii) because in adabatic process no heat change occurs.

S3. The correct answer is (ii)

S4. The correct answer is (iii). The combustion of methane is represented on

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

Here, $\Delta n_{(g)} = 1 - 3 = -2$

Thus, $\Delta H^{\ominus} < \Delta U^{\ominus}$.

S5. The correct answer is (iv). Since the reaction is exothermic ($\Delta H < 0$) and has $\Delta S > 0$, therefore, it is feasible at all temperatures because both the factors favour spontaneity.

S6. ΔH : Negative (–ve) because energy is released in bond formation.

ΔS : Negative (–ve) because entropy decreases when atoms combine to form molecules.

S7. For $NO(g)$; $\Delta_f H^{\ominus}$ = +ve: Unstable in nature.

For $NO_2(g)$; $\Delta_f H^{\ominus}$ = –ve: Stable in nature.

S8. (a) When a gas expands against the external pressure p_{ex} in an irreversible manner, work is done by the gas and it is given by the expression; $w_{PV} = p_{ex}\Delta V$

(b) When gas compressed, work is done on the gas.

(c) When gas expands into vacuum, on work is done because external pressure is zero.

(d) In case the expansion is done under reversible conditions, the work done by the gas is maximum because the opposing force is infinitessimally smaller than the driving force.

S9. The instant packs consist of a plastic bags containing a pouch of water and a suitable chemical substance in dry state. In cold packs, the chemical substance used is generally ammonium nitrate whose dissolution process is endothermic. As a result it is gives cold sensation.

$$NH_4NO_3 \xrightarrow{\text{Water}} NH_4^+ + NO_3^-; \Delta H = +26.2 \text{ kJ}$$

In hot instant pack, the chemical compound used is $MgSO_4$ or $CaCl_2$, the dissolution of which is exothermic. As a result it gives hot sensation,

$$CaCl_2(s) \xrightarrow{\text{Water}} Ca^{2+}(aq) + 2Cl^-(aq); \Delta H = -82.2 \text{ kJ}$$

These packs are generally used by athletes as first aid devices for treatment of injuries. Their effect lasts for about 15–20 minutes.

S10. (i) $\Delta_r H$ refers to heat change taking place at constant temperature and constant pressure. Since most of the reactions that we carry out in laboratory occur in open vessels *i.e.*, at the atmospheric pressure. Hence, it is more convenient to determine the value of $\Delta_r H$ as compared to $\Delta_r U$.

(ii) Enthalpy change depends upon the conditions of the reaction. Thus, in order to make a comparative study, it is necessary to chose same standard reference state of various substances, and express the value of $\Delta_f H$ at the standard state.

(iii) Physical changes like fusion, vaporisation, sublimation, etc., and phase transitions such as diamond to graphite or S_8 (Monoclinic) are also accmpanied by the change of enthalpy. Therefore, while writing the thermochemical equation, it is necessary to specify to physical state or phase of the substance involved.

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Q1. For the following equilibrium, $K_c = 6.3 \times 10^{14}$ at 1000 K.

$$NO(g) + O_3(g) \rightleftharpoons NO_2(g) + O_2(g)$$

Both the forward and reverse reactions in the equilibrium are elementary bimolecular reactions. What is K_c, for the reverse reaction?

Q2. Calculate the value of K_p at 100 °C for the reaction $N_2O_4(g) \rightleftharpoons 2NO_2(g)$. It is given that at equilibrium partial pressures of N_2O_4 and NO_2 are 0.002 bar and 0.4 bar respectively at 100 °C.

Q3. What is K_c for the following equilibrium when the equilibrium concentration of each substance is: $[SO_2] = 0.60$ M, $[O_2] = 0.82$ M and $[SO_3] = 1.90$ M?

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Q4. At 900 K a sealed bulb contains equilibrium mixture of SO_3 at 0.23 atm, SO_2 at 0.15 atm and O_2 at 0.73 atm. What is the equilibrium constant for $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$?

Q5. Equilibrium constant, K_c for the reaction, $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at 500 K is 0.061. At a particular time, the analysis shows that composition of the reaction mixture is 3.00 mol L^{-1} N_2, 200 mol L^{-1} H_2, and 0.500 mol L^{-1} NH^3. Is the reaction at equilibrium? If not, which direction does the reaction tend to proceed to reach equilibrium?

Q6. For the reaction $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ $K_p = 0.157$ atm at 300 K. Calculate the value of K_c for the same reaction at the same temperature.

Q7. For the reaction $2A(g) \rightarrow B(g) + 2C(g)$ the equilibrium constant K_p is 1.776×10^{-5} atm at 700 K. Calculate K_c for the reaction.

Q8. The dissociation constant for the reaction

$$CuO(s) \rightleftharpoons Cu(s) + \frac{1}{2}O_2(g) \text{ is } 2 \times 10^{-22}(atm)^{1/2}$$

What is the pressure of oxygen?

Q9. At 298 K, the equilibrium constant K_c is 2.0×10^{15} for the reaction

$$Cu(s) + 2Ag^+(aq) \rightleftharpoons Cu^{2+}(aq) + 2Ag(s)$$

If at equilibrium $[Ag^+] = 1.0 \times 10^{-9}$ mol L^{-1}, what is the equilibrium concentration of $[Cu^{2+}]$?

Q10. For the reaction, $Cu(s) + 2Ag^+(aq) \rightleftharpoons Cu^{2+}(aq) + 2Ag(s)$, the equilibium constant is 2.0×10^{15} L mol^{-1} at 278 K. In a solution copper has displaced some silver ions so that the concentration of Ag+ ions is 3.0×10^{-9} mol L^{-1} and the concentration of Cu^{2+} ions is 1.8 10^{-2} mol^{-1}. Is this system at equilibrium?

Q11. The equilibrium constant for the decomposition of solid ammonium chloride into gaseous anmonia and hydrogen chloride is 6×10^{-9} atm^2 at 400 K. Calculate the equilibrium pressure of the two gases.

Q12. For the reaction $Cu(s) + 2Ag^+(aq) \rightleftharpoons Cu^{2+}(aq) + 2Ag(s)$ $K = 2 \times 10^{15}$ at 298 K. Calculate the equilibrium constant for

$$\frac{1}{2}Cu(s) + Ag^+(aq) \rightleftharpoons \frac{1}{2}Cu^{2+}(aq) + Ag(s) \quad \text{... (i)}$$

Q13. At 298 K the equilibrium constant is 138 for the reaction

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftarrows FeSCN^{2+}(aq) \quad \text{... (A)}$$

Calculate the equilibrium constant at 298 K for the following reactions:

(a) $2Fe^{3+}(aq) + 2SCN^{-}(aq) \rightleftarrows 2FeSCN^{2+}(aq)$

(b) $FeSCN^{2+}(aq) \rightleftarrows Fe^{3+}(aq) + SCN^{-}(aq)$

Q14. For the reaction $PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g)$ the degree of dissociation of PCl_5 is 82% at 520 K and 1 bar. Calculate the equilibrium constant K_p.

Q15. For the reaction $N_2O_5(g) \rightleftarrows 2NO_2(g) + 0.5O_2(g)$, calculate the mole fraction of $N_2O_5(g)$ decomposed at a constant volume and temperature, if the initial pressure is 600 mm Hg and the pressure at any time is 960 mm Hg. Assume ideal gas behabiour.

Q16. The degree of dissociation is 0.4 at 400 K and 1.0 atm for the gaseous reaction $PCl_5 \rightleftarrows PCl_3 + Cl_2$. Assuming ideal behaviour of all the gases, calculate the density of equilibrium mixture at 400 K and 1.0 atmosphere. (Atomic mass of P = 31.0 and Cl = 35.5)

Q17. At 700 K equilibrium constant for the reaction $H_2(g) + I_2(g) \rightleftarrows 2HI(g)$ is 54.8. If 0.5 mol L^{-1} of HI (*g*) is present at equilibrium what are the concentrations of $H_2(g)$ and $I_2(g)$ assuming that we initially started with HI (*g*) and allowed it to reach equilibrium at 700 K.

Q18. A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was attained, concentration of PCl_5 was found to be 0.05 mol L^{-1}. If the value of K_c is 8.3×10^3 mol L^{-1}, what are the concentrations of PCl_3 and Cl_2 at equilibrium?

Q19. The equilibrim constant for the reaction $H_2(g) + Br_2(g) \rightleftarrows 2HBr(g)$ is 1.6×10^5 at 1024 K. Find the equilibrium pressure of all gases if HBr at 10 bar is introduced into a sealed container at 1024 K.

Q20. At a certain temperature and a total pressure of 10^5 kPa, iodine vapour contains 40% by volume of iodine atoms. Calculate K_p for the equilibrium reaction $I_2(g) \rightleftarrows 2I(g)$.

Q21. At 473 K, equilibrium constant, K_c for decomposition of phosphorus pentachloride, PCl_5 is 8.3×10^{-3}. If decompositionis depicted as $PCl_5(s) \rightleftarrows PCl_3(g) + Cl_2(g)$; $\Delta H = 124.0$ kJ mol^{-1}.

(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 PCl_5 is added (ii) the pressure is increased (iii) the temperature is increased?

Q22. A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was attained, the concentration of PCl_5 was found to be 0.05 mol L^{-1}. What are the concentrations of PCl_3 and Cl_2 at equilibrium if the value of K_c is 8.3×10^{-3} mol/L?

Q23. A sample of HI(*g*) is placed in flask at a pressure of 0.2 atm. At equilibrium, the partial pressure of HI(*g*) is 0.04 atm. What is K_p for the given equilibrium?

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$

Q24. At 450 K, $K_p = 2.0 \times 10^{10}$/bar for the given reaction at equilibrium

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

What is K_c at this temperature?

Q25. A mixture of 1.57 mol of N_2, 192 mol of H_2 and 8.13 mol of NH_3 is introduced into a 20 L reaction vessel at 500 K. At this temperature, the equilibrium constant, K_c for the reaction

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \text{ is } 1.7 \times 10^2.$$

Is the reaction mixture at equilibrium? If not, what is the direction of th net reaction?

Q26. What is the equilibrium concentration of each of the substance in the equilibrium when the initial concentration of ICl was 0.78 M?

$$2ICl(g) \rightleftharpoons I_2(g) + Cl_2(g); \quad K_c = 0.14$$

Q27. One mole of H_2O and one mole of CO are taken in 10 L vessel and heated to 725 K. At equilibrium 40% of water (by mass) reacts with CO according to the equation,

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

Calculate the equilibrium constant for the reaction.

Q28. A mixture of SO_3, SO_2 and O_2 gases is maintained in a 10 L flask at a temperature at which the equilibrium constant (K_c) is 100 for the reaction $2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g)$:

(a) If the amounts of SO_2 and SO_3 in the flask are eual, how many moles of O_2 are present?

(b) If the amount of SO_3 in the flask is twice the amount of SO_2, how many moles of O_2 are present?

Q29. At 700 K the equilibrium constant K_p for the reaction

$$2SO_3(g) \rightleftarrows 2SO_2(g) + O_2(g)$$

is 1.80×10^{-3} k Pa. Calculate the value of K_c in mol L^{-1} for the same reaction at 700 K.

Q30. For the reaction, $CH_4(g) + 2H_2S(g) \rightleftarrows CS_2(g) + 4H_2(g)$, at value of the equilibrium constant, K_c is 3.6. For each of the following composition, decide whether reaction mixture is a equilibrium. If it is not, decide to which direction the reaction should go.

(a) $[CH_4]$ = 10.7 M, $[H_2S]$ = 1.20 M, $[CS_2]$ = 0.90 M, $[H_2]$ = 1.78 M

(b) $[CH_4]$ = 1.45 M, $[H_2S]$ = 1.29 M, $[CS_2]$ = 1.25 M, $[H_2]$ = 1.75 M

Q31. Find out the value of K_c for each of the following equilibria from the value of K_p

(a) $2NOCl(g) \rightleftarrows 2NO(g) + Cl_2(g)$ $\quad K_p = 1.8 \times 10^{-2}$ bar at 500 K

(b) $CaCO_3(s) \rightleftarrows CaO(s) + CO_2(g)$ $\quad K_p = 167$ bar at 1073 K

Q32. Prove that the equilibrium constants for the two reactions

(a) $H_2(g) + I_2(g) \rightleftarrows 2HI(g)$ $\quad$ (b) $2HI(g) \rightleftarrows H_2(g) + I_2(g)$

are related by $K_p(ii) = 1/K_p(i)$.

Q33. For the reaction

$$SO_2(g) + \frac{1}{2}O_2(g) \rightleftarrows SO_3(g)$$

the equilibrium constant $K_c = 1.7 \times 10^{12}$ at 300 K. Calculate the equilibrium constants for the following reactions at 300 K:

(a) $SO_2(g) + \frac{1}{2}O_2(g) \rightleftarrows 2SO_3(g)$ $\quad$ (b) $SO_3(g) \rightleftarrows SO_2(g) + \frac{1}{2}O_2(g)$.

Q34. For the reaction $N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$ if $K_p = 41$ at 400 K what is the value of K_p for each of the following reactions at the same temperature?

(a) $2NH_3(g) \rightleftarrows N_2(g) + 3H_2(g)$ $\quad$ (b) $\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightleftarrows NH_3(g)$

(c) $2N_2(g) + 6H_2(g) \rightleftarrows 4NH_3(g)$

Q35. Nitric oxide reacts with bromine and gives nitrosyl-bromide as per reaction 2 NO (*g*) + Br_2 (*g*) ⇄ 2 NOBr (*g*). When 0.087 mol of NO and 0.0437 mol of Br_2 are mixed in a closed container at constant temperature, 0.0518 mol of NOBr is obtained at equilibrium. Calculate equilibrium amount of nitric oxide and bromine. What is the value of K_c?

Q36. Bromine monochloride (BrCl) decomposes into bromine and chlorine and reaches the equilibrium 2BrCl(*g*) ⇄ Br_2(*g*) + Cl_2(*g*) for which K_c = 32 at 500 K. If initially pure BrCl is present at a concentration of 3.30×10^{-3} mol L^{-1} what is its molar concentration in the mixture at equilibrium?

Q37. A mixture of 1 mole of H_2O and 2 mole of CO is taken in a 10 litre container and heated to 725 K. At equilibrium 40 per cent of water by mass reacts with carbon monoxide according to the equation

$$CO\,(g) + H_2O\,(g) \rightleftarrows CO_2\,(g) + H_2\,(g)$$

Calculate the equilibrium constants K_c for the given reaction.

Q38. The composition of the equilibrium mixture ($Cl_2 \rightleftarrows 2Cl$), which is attained at 1200 °C, is determined by measuring the rate of effusion through a pin-hole. It is observed that at 1.80 mm Hg pressure, the mixture effuses 1.16 times as fast as krypton effuses under the same conditions. Calculate the fraction of chloride molecules dissociated into atoms, (Relative atomic mass of Kr = 84)

Q39. At 1127 K and 1 atm pressure, a gaseous mixure of CO and CO_2 in equilibrium with solid carbon has 90.55% CO by mass

$$C(s) + CO_2(g) \rightleftharpoons 2CO(g)$$

Calculate K_c for this reaction at the above temperatue.

Q40. Reaction between N_2 and O_2 takes place as follows:

$$2N_2(g) + O_2(g) \rightleftharpoons 2N_2O(g)$$

If a mixture of 0.482 mol N_2 and 0.933 mol of O_2 is placed in a 10 L reaction vessel and allowed to form N_2O at a temperature for which $K_c = 2.0\ 10^{-27}$, determine the composition of equilibrium mixture.

Q41. One of the reactions that takes place in producing steel from iron ore is the reduction of iron (II) oxide by carbon monoxide to given iron metal and CO_2.

$$FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g);\quad K_p = 0.265 \text{ atm at 1050 K.}$$

What are the equilibrium partial pressures of CO and CO_2 at 1050 K if the initial pressures are:

$$p_{CO} = 1.4 \text{ atm} \quad \text{and} \quad p_{CO_2} = 0.80 \text{ atm?}$$

Q42. Three moles of PCl_5 are placed in a 100 L vessel which contains 1 mole N_2 and is maintained at 500 K. The equilibrium pressure is 2.05 atm. Calculate the equilibrium constant Kp for the dissociation of PCl_5. Take R = 0.0820 atm L mol^{-1} K^{-1}.

Q43. In the reaction representing the dissociation of phosgene gas ($COCl_2$)

$$COCl_2\,(g) \rightleftarrows CO\,(g) + Cl_2\,(g)$$

the pressure at equilibrium is increased, that is the mixture of gases is compressed keeping the temperature constant. What is the effect of this compression on (a) the eqilibrium concentration of CO, (b) the partial pressure of $COCl_2$, and (c) the equilibrium constant (K_p) of the reaction?

Q44. Two moles of HI are confined in one litre vessel at 760 K. The dissociation reaction takes place as $2HI(g) \rightleftarrows H_2(g) + I_2(g)$

For the reaction $K_c = 0.023$ at 760 K. Compute the equilibrium concentrations of H_2, I_2 and HI. What is the degree of dissociation of HI at 760 K?

Q45. One mole of $N_2O_4(g)$ is confined in a 24.6 litre vessel at 300 K. The dissociation reaction takes place as $N_2O_4(g) \rightleftarrows 2NO_2(g)$.

For the reaction, $Kc = 4.6 \times 10^{-3}$ calculate

(a) the degree of dissociation of N_2O_4, and (b) the equilibrim concentration of NO_2.

Q46. Expression for K_c for the reaction $N_2O_4(g) \rightleftarrows 2NO_2(g)$

Q47. Expression for K_c for the $PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g)$

Q48. Calculate $\Delta_r G^{\ominus}$ and K for the reaction $NO(g) + 0.5\,O_2(g) \rightleftarrows NO_2(g)$ at 298 K. Given that

Gas	NO_2	NO	O_2
$\Delta_f G^{\ominus}$/kJ mol	52.0	87.0	0.0

Q49. Ethyl acetate is formed by the reaction of ethanol and acetic acid. The equilibrium is represented by

$$CH_3COOH(l) + C_2H_5OH(l) \rightleftarrows CH_3COOC_2H_5(l) + H_2O(l)$$

(a) Name the reactants and products in this reaction.

(b) Express the concentration ratio for the reaction.

(c) At 293 K, when we start with 1.0 mole of acetic acid and 0.18 mole of ethanol 0.171 mole of ethyl acetate is formed at equilibrium. Calculate the equilibrium constant.

(d) Starting with 0.50 mole of ethanol and 1.00 mole of acetic acid 0.214 mole of ethyl acetate is formed after some time at 293 K. Has equilibrium been reached?

(e) Why is some concentrated sulphuric acid added to the reaction mixture in the laboratory preparation of ethyl acetate?

(f) We do not use dilute sulphuric acid for this reaction. Give reason.

(g) The heat of reaction is nearly zero for the given reaction. Tell how will the equilibrium constant depend upon the temperature?

Q50. At 288 K the solubility of iodine in water is 1.1×10^{-3} mol/L. In an experiment 200 mg of iodine is strirred in 100 mL of water till equilibrium ia reached.

(a) Calculate the mass of iodine in the resultiong saturated solution at 288 K.

(b) What will be the mass of the iodine that is left undissolved?

(c) After equilibrium is reached. 150 mL of water is added to the above saturated solution and stirred will till new equilibrium is established.

(i) How much iodine will be dissolved in the final solution?

(ii) How much iodine will be left undissolved?

(iii) What will be the concentration of iodine solution?

(*Atomic mass of iodine = 127u*)

Q51. For the reaction $N_2O_4(g) \rightleftarrows 2NO_2(g)$ the equilibrium constant is 80 atm at 100 °C. Predict the behaviour of the following mixtures of NO_2 and N_2O_4.

(a) $p(NO_2) = p(N_2O_4) = 4 \times 10^{-3}$ atm (b) $p(N_2O_4) = 2 \times 10^{-3}$ atm and $p(NO_2) = 0.4$ atm

(c) $p(N_2O_4) = 2 \times 10^{-3}$ atm and $p(NO_2) = 1$ atm

(d) $p(N_2O_4) = 0$ and $p(NO_2) = 2 \times 10^{-3}$ atm

Q52. For the reaction $H_2(g) + I_2(g) \rightleftarrows 2HI(g)$, the equilibrium constant is 0.35 at 298 K. Predict the direction of the reaction in the following mixtures.

(a) Partial pressure of $H_2 = p(H_2) = 0.10$ atm, and $p(HI) = 0.80$ atm and there is solid I_2 in the container.

(b) $p(H_2) = 0.55$ atm, $p(HI) = 0.44$ atm, and there is solid I_2 in the container.

(c) $p(H_2) = 2.5$ atm, $p(HI) = 0.15$ atm, and there is solid I_2 in the container.

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S1.

$$K'_c = \frac{1}{6.3 \times 10^{14}} = \mathbf{1.59 \times 10^{-15}}.$$

S2. For the reaction $N_2O_4(g) \rightleftarrows 2NO_2(g)$.

$$K_p = \frac{(pNO_2)^2}{pN_2O_4} = \frac{(0.4\ \text{bar})^2}{0.002\ \text{bar}} = \mathbf{80\ bar}.$$

S3.

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

$$= \frac{(1.90)^2}{(0.60)^2(0.82)} = \mathbf{12.33\ mol\ L^{-1}}.$$

S4. For the reaction $2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g)$

$$K_p = \frac{p^2(SO_3)}{p^2(SO_2) \times p(O_2)}$$

The partial pressures of gases are given

$$p(SO_3) = 0.23\ \text{atm}$$
$$p(SO_2) = 0.15\ \text{atm}$$
$$p(O_2) = 0.73\ \text{atm}$$

$$K_p = \frac{(0.23\ \text{atm})^2}{(0.15\ \text{atm})^2 \times (0.73\ \text{atm})}$$
$$= 3.22\ \text{atm}^{-1}.$$

S5. For the reaction $N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$ $K_c = 0.061$, and at a particular time

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(0.500)^2}{(3.00) \times (2.00)^3}$$

$$= \frac{0.25}{24} = 0.0104 < 0.061$$

(a) Since $Q_c < K_c$, the reaction is not at equilibrium.

(b) The reaction will proceed in the forward direction to form more NH_3 to reach equilibrium.

S6. For a reaction K_p and K_c are related by

$$K_p = K_c(RT)^{\Delta v}$$

For the reaction $N_2O_4(g) \rightleftarrows 2NO_2(g)$

$$\Delta v = 2 - 1 = 1$$

$$K_p = 0.157 \text{ atm at } T = 300 \text{ K}$$

Now, $R = 0.082 \text{ atm L mol}^{-1} \text{ K}^{-1}$

$$\therefore \quad K_c = \frac{K_p}{(RT)^{\Delta v}} = \frac{0.157 \text{ atm}}{0.082 \text{ atm L mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}}$$

$$= 6.38 \times 10^{-3} \text{ mol L}^{-1}.$$

S7. We know that, $K_p = K_c (RT)^{\Delta v}$

$$\therefore \quad K_c = \frac{K_p}{(RT)^{\Delta v}} \quad \dots \text{(i)}$$

Now, $K_p = 1.776 \times 10^{-5}$ atm, $R = 0.082$ atm L mol^{-1} K^{-1}, $T = 700$ K. For the reaction

$$2A(g) \rightleftarrows B(g) + 2C(g)$$

$$\Delta v = 1 + 2 - 2 = 3 - 2 = 1$$

On substituting the values of known quantities into equation (i), we get

$$K_c = \frac{1.776 \times 10^{-5} \text{ atm}}{0.082 \text{ atm L mol}^{-1} \text{K}^{-1} \times 700 \text{ K}}$$

$$= 3.09 \times 10^{-7} \text{ mol L}^{-1}.$$

S8. For the heterogeneous reaction,

$$CuO(s) \rightleftarrows Cu(s) + \frac{1}{2} O_2(g)$$

The ewuilibrium constant is given in terms of the pressure of O_2 by

$$K_p = (pO_2)^{1/}2$$

$$\therefore \quad p(O_2) = K_p^2$$

$$= (2 \times 10^{-22} \text{ atm}^{1/2})^2$$

$$= 4 \times 10^{-44} \text{ atm}$$

Partial pressure of O_2 is 4×10^{-44} atm.

S9. For the reaction $Cu(s) + 2Ag^+(aq) \rightleftarrows Cu^{2+}(aq) + 2Ag(s)$

$$K_c = \frac{[Cu^{2+}]}{[Ag^+]^2} \quad \dots \text{(i)}$$

On substituting the given values of

$$K_c = 2.0 \times 10^{15}\ \text{L mol}^{-1},$$
$$[Ag^+] = 1.0 \times 10^{-9}\ \text{mol L}^{-1}$$

In equation (i), we have

$$2.0 \times 10^{15}\ \text{L mol}^{-1} = \frac{[Cu^{2+}]}{(1.0 \times 10^{-9}\ \text{mol L}^{-1})^2}$$

$$\therefore \quad [Cu^{2+}] = 2.0 \times 10^{15}\ \text{L mol}^{-1} \times 1.0 \times 10^{-18}\ \text{mol}^2\ \text{L}^{-2}$$
$$= 2.0 \times 10^{-3}\ \text{mol L}^{-1}.$$

S10. For the reaction, $Cu(s) + 2Ag^+(aq) \rightleftarrows Cu^{2+}(aq) + 2Ag(s)$, the quotient ($Q_c$) is given by:

$$Q_c = \frac{[Cu^{2+}]}{[Ag^+]^2} = \frac{1.8 \times 10^{-2}\ \text{mol L}^{-1}}{(3.0 \times 10^{-9}\ \text{mol L}^{-1})^2} = 2.0 \times 10^{15}.$$

Since quotient (Q_c) for the given reaction is equal to its equilibrium constant (K_c) the system tis at equlibrium.

S11. The decomposition of solid ammonium chloride is representd by

$$NH_4Cl(s) \rightleftarrows NH_3(g) + HCl(g)$$

For this heterogeneous reaction the equilibrium constant is expressed as

$$K_p = p\,(NH_3) \times p\,(HCl)$$

Since NH_3 and HCl are formed in equal amounts, their pressures are also same at equilibrium. Thus $p\,(NH_3) = p\,(HCl)$ and

$$K_p = p\,(NH_3) \times p\,(HCl) = p^2(NH_3) = p^2(HCl)$$

$$\therefore \quad p\,(NH_3) = p\,(HCl) = \sqrt{K_p} = \sqrt{6 \times 10^{-9}\ \text{atm}^2}$$
$$= 7.7 \times 10^{-5}\ \text{atm}.$$

S12. Equation (i) is obtained when equation (A) is multiplied by half. Thus:

$$K(i) = [K(A)]^{1/2} = (2 \times 10^{15})^{1/2}$$
$$= (20 \times 10^{14})^{1/2} = 4.47 \times 10^7.$$

S13. (a) When equation (A) is multiplied by 2, we get equation (i). Therefore:

$$K(i) = \{K(A)\}^2 = (138)^2 = 19044 = 1.9 \times 10^4$$

(b) Equation (ii) is obtained when equation (A) is reversed. Thus:

$$K(ii) = \frac{1}{K(A)} = \frac{1}{138} = 7.2 \times 10^{-3}.$$

S14. It is given that α = 82% = 82/100 = 0.82 and P = 1 bar

$$\therefore \quad K_p = \frac{\alpha^2 P}{1-\alpha^2} = \frac{(0.82) \times 1}{1-(0.82)^2}\ \text{bar} = 2.052\ \text{bar}.$$

S15. $N_2O_5(g) \rightarrow 2NO_2(g) + 0.5O_2(g)$

	$N_2O_5(g)$	$2NO_2(g)$	$0.5O_2(g)$
Initial	P_0	0	0
When $t = t$	$P_0(1-\alpha)$	$2P_0\alpha$	$0.5\,P_0\alpha$

$$P\,(\text{total}) = P_0(1 - \alpha + 2\alpha + 0.5\alpha) = P_0(1 + 1.5\alpha)$$

$\therefore$ 960 mm Hg = 600 mm Hg $(1 + 1.5\alpha)$

or, $\alpha = 0.4$ = Mole fraction of N_2O_5 decomposed.

S16. Let us suppose that n is the strating amount of PCl_5. That is $n = m/M = m/208.5 \text{ g mol}^{-1}$.

For the reaction:	PCl_5	$\rightleftarrows$	PCl_3	+	Cl_2
Initial amounts	n		0		0
Equilibrium amounts	$n(1-\alpha)$		$n\alpha$	+	$n\alpha$

Total amount of mixture at equilibrium = $n(1-\alpha) + n\alpha + n\alpha$

$= n(1+\alpha) = n \times 1.4 = 1.4 \times \text{m/M}$

For ideal gas

$$PV = n_{\text{mixture}} \times RT = (n \times 1.4)\,RT = \frac{m}{M} \times 1.4 \times \text{m/M}$$

$$\therefore \quad \frac{m}{V} = \frac{PM}{1.4\,RT} = \frac{1\,\text{atm} \times 208\text{ g mol}^{-1}}{1.4 \times 0.082\text{ atm L mol}^{-1}K^{-1} \times 400\,K} = 4.53\text{ g/L}$$

Thus, density of equilibrium mixture is 4.53 g/L.

S17. For the reaction $H_2(g) + I_2(g) \rightleftarrows 2HI(g)$

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{[HI]^2}{[H_2]^2} = \frac{[HI]^2}{[I_2]^2}$$

$$\therefore \quad \frac{[HI]}{[H_2]} = (K_c)^{1/2}$$

and $$[H_2] = [I_2] = \frac{[HI]}{(K_c)^{1/2}} = \frac{0.5\text{ mol L}^{-1}}{(54.8)^{1/2}}$$

$$= 0.0675\text{ mol L}^{-1}.$$

S18. For the traction $PCl_5(g) \longrightarrow PCl_3(g) + Cl_2(g)$

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{x^2}{0.05\text{ mol L}^{-1}}$$

$$\therefore \quad [PCl_3] = [Cl_2] = x = (K_c \times 0.05\text{ mol L}^{-1})^{1/2}$$

$$= (8.3 \times 10^3\text{ mol L}^{-1} \times 0.05\text{ mol L}^{-1})^{1/2}$$

$$= 2.04\text{ mol L}^{-1}.$$

S19. Let us suppose that x is the equilibrium pressure of each one of H_2 and Br_2, and p (= 10 bar) is that of HBr. Thus, for the reaction $H_2(g) + Br_2(g) \rightleftarrows 2HBr(g)$

$$K_p = \frac{(p\text{HBr})^2}{(pH_2) \times (pBr_2)} = \frac{p^2}{x^2}$$

and $$x^2 = \frac{p^2}{K_p} = \frac{(10 \text{ bar})^2}{1.6 \times 10^5} = 6.25 \times 10^{-4} \text{ bar}^2$$

$$x = (6.25 \times 10^{-4} \text{ bar}^2)^{1/2} = 2.5 \times 10^{-2} \text{ bar}$$
$$x = p(H_2) = p(Br^2) = 2.5 \times 10^{-2} \text{ bar}$$
$$p = (\text{HBr}) = 10 \text{ bar.}$$

S20. For the reaction $I_2(g) \rightleftarrows 2I(g)$

$$K_p = (pI)^2/pI_2 \qquad \text{... Dalton's law}$$

Partial pressure = Mole fraction × Total pressure

Mole fraction of gas = Volume % if gas.

Thus Mole fraction of I = 40/100 = 0.4

Mole fraction of I_2 = 1 – 0.4 = 0.6

Partial pressure of I = $pI = 0.4 \times 10^5$ k Pa

Partial pressure of I_2 = $pI_2 = 0.6 \times 10^5$ k Pa

$$K_p = \frac{(pI)^2}{pI_2} = \frac{(0.4 \times 10^5 \text{ kPa})^2}{0.6 \times 10^5 \text{ kPa}}$$
$$= 2.67 \times 10^4 \text{ kPa.}$$

S21. (a) For the reaction $PCl_5(s) \rightleftarrows PCl_3(g) + Cl_2(g)$

$$K_c(a) = [PCl_3][Cl_2] = 8.3 \times 10^{-3}$$

(b) For the reverse reaction $PCl_3(g) + Cl_2(g) \rightleftarrows PCl_5(g)$

$$K_c(b) = \frac{1}{[PCl_3][Cl_2]} = \frac{1}{K_c(a)} = \frac{1}{8.3 \times 10^{-3}} = 120.48$$

(c) ***Change of K_c***

It is independent of the initial amount of the reactant and also independent of total pressure. Therefore,

(i) Addition of more PCl_5 will not affect the value of K_c

(ii) Increase of pressure will not affect the value of K_c

(iii) Increase of temperature will increase the value of K_c (reaction is endothermic).

S22. For the reaction $PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g)$

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

Or, $[PCl_3] \times [Cl_2] = [Cl_2]^2 = [PCl_3]^2 = K_c \times [PCl_5] = 8.3 \times 10^{-3} \times 0.05$ mol/L

$\therefore$ $[PCl_3] = [Cl_2] = (8.3 \times 10^{-3} \text{ mol/L} \times 0.05 \text{ mol/L})^{1/2} = 0.02$ mol/L

S23.

	$2HI(g)$ $\rightleftharpoons$	$H_2(g)$ +	$I_2(g)$
Initial pressure	0.2 atm	—	—
At equilibrium	$(0.2 - p)$ atm	$p/2$ atm	$p/2$ atm
	(0.04 atm)	0.08 atm	0.08 atm

$$K_p = \frac{P_{H_2} \times P_{I_2}}{(P_{HI})^2} = \frac{0.08 \times 0.08}{(0.04)^2} = \mathbf{4}.$$

S24. For the given reaction,

$$\Delta n_g = n_p - n_r = 2 - 3 = -1$$

$$Kp = K_c(RT)^{\Delta n}$$

or $$Kc = K_p(RT)^{-\Delta n} = K_p(RT)$$

$$= (2.0 \times 10^{10} \text{ bar}^{-1})(0.0831) \text{ L bar K}^{-1} \text{ mol}^{-1})(450 \text{ K})$$

$$= \mathbf{5.48 \times 10^{11} \text{ L mol}^{-1}}.$$

S25. $$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$= \frac{(8.13/20 \text{ mol L}^{-1})^2}{(1.57/20 \text{ mol L}^{-1})(1.92/20 \text{ mol L}^{-1})^3}$$

$$= 2.38 \times 10^3.$$

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

Since $Q_c > K_c$, the net reaction will be in the backward direction.

S26. Suppose at equilibrium,

$$[I_2] = [Cl_2] = x \text{ mol L}^{-1}$$

	$2ICl(g)$ $\rightleftharpoons$	$I_2(g)$ +	$Cl_2(g)$
Initial conc.	0.78 M	0	0
At equilibrium	$0.78 - 2x$	x	x

$$K_c = \frac{[I_2][Cl_2]}{[ICl]}$$

$\therefore$ $$0.14 = \frac{x \times x}{(0.78 - 2x)^2}$$

or $$\frac{x}{0.78-2x} = \sqrt{0.14} = 0.37$$

or $$x = 0.29 - 74\,x$$

or $$1.74 = 0.29 \quad \text{or} \quad x = 0.17$$

Hence, at equilibrium

$$[I_2] = [Cl_2] = \mathbf{0.17\ M},$$
$$[ICl] = 0.78 - 2 \times 0.17\,M = \mathbf{0.44\,M}.$$

S27. At equilibrium

$$[H_2O] = \frac{1-0.40}{10}\ \text{mol L}^{-1} = 0.06\ \text{mol L}^{-1}$$
$$[CO] = 0.06\ \text{mol L}^{-1}$$

$$[H_2] = \frac{0.40}{10}\ \text{mol L}^{-1} = 0.04\ \text{mol L}^{-1}$$
$$[CO_2] = 0.04\ \text{mol L}^{-1}$$

$$K = \frac{[H_2][CO_2]}{[H_2O][CO]} = \frac{0.04 \times 0.04}{0.06 \times 0.06} = \mathbf{0.44}.$$

S28. Equilibrium state is described as:

$$2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g).$$

$$K_c = \frac{[SO_3]^2}{[SO_2]^2\,[O_2]}$$

$\therefore$ $$[O_2] = \frac{[SO_3]^2}{[SO_2]^2} \times \frac{1}{K_c} = \text{Molarity of } O_2.$$

(a) $\because$ $[SO_3] = [SO_2]$ ... [Given]

$\therefore$ $$[O_2] = 1 \times \frac{1}{100} = 0.01\ \text{mol L}^{-1}$$

and amount of O_2 = Molarity $\times$ Volume
$= 0.01\ \text{mol L}^{-1} \times 10\ \text{L} = 0.1$ mol.

(b) $\because$ $[SO_3] = 2\,[SO_2]$ ... [Given]

$\therefore$ $$[O_2] = \frac{[SO_3]^2}{[SO_2]^2} \times \frac{1}{K_c} = \frac{[2\,SO_2]^2}{[SO_2]^2} \times \frac{1}{100} = 0.04\ \text{mol L}^{-1}$$

and amount of O_2 = Molarity $\times$ Volume
$= 0.04\ \text{mol L}^{-1} \times 10\ \text{L} = 0.4$ mol.

S29. We know that $$K_p = K_c\,(RT)^{\Delta v}$$

$\therefore$ $$K_c = \frac{K_p}{(RT)^{\Delta v}} \quad \text{... (i)}$$

It is given that $K_p = 1.80 \times 10^{-3}$ kPa = $1.80 \times 10^{-3} \times 10^3$ Pa = 1.80 Pa = 1.80 N m^{-2}. Since K_p is in N m^{-2}, we shall use $R = 8.314$ N m mol^{-1} K^{-1} and $T = 700$ K.

For the reaction $2SO_3(g) \rightleftarrows 2SO_2(g) + O_2(g)$

$$\Delta v = (2 + 1) - 2 = 3 - 2 = 1$$

On substituting the values of known quantities into equation (i), we get

$$K_c = \frac{1.80\ \text{N m}^{-2}}{8.314\ \text{N m mol}^{-1}\text{K}^{-1} \times 700\ \text{K}}$$

$$= 3.09 \times 10^{-4}\ \text{mol m}^{-3}.$$

$$= 3.09 \times 10^{-7}\ \text{mol L}^{-1}. \qquad (\because\ 10^{-3}\ \text{m}^3 = 1\ \text{dm}^3 = 1\ \text{L})$$

S30. For the reaction $CH_4(g) + 2H_2S(g) \rightleftarrows CS_2(g) + 4H_2(g)$ $K_c = 3.6$

(a) $$Q_c = \frac{[CS_2][H_2]^2}{[CH_4][H_2S]^2} = \frac{(0.90\ \text{M}) \times (1.78\ \text{M})^4}{(1.07\ \text{M}) \times (1.20\ \text{M})^2} = 6.51 > 3.6$$

$\because$ $Q_c > K_c$, therefore the reaction mixture is not at equilibrium. The reaction will proceed in the reverse direction to form more reactants to reach equilibrium.

(b) $$Q_c = \frac{[CS_2][H_2]^2}{[CH_4][H_2S]^2} = \frac{(1.25) \times (1.75)^4}{(1.45) \times (1.29)^2} = 3.88 > 3.6$$

$\because$ $Q_c > K_c$, therefore the reaction mixture is not at equilibrium. It will proceed in the reverse direction to form more reactants to reach equilibrium.

S31. The reaction $K_p = K_c(RT)^{\Delta v}$ gives

$$K_c = \frac{K_p}{(RT)^{\Delta v}}$$

(a) For the reaction $2NOCl(g) \rightleftarrows 2NO(g) + Cl_2(g)$

$$\Delta v = (2 + 1) - 2 = 1, \quad K_p = 1.8 \times 10^{-2}\ \text{bar at 500 K}$$

$\therefore$ $$K_c = \frac{K_p}{(RT)^{\Delta v}} = \frac{1.8 \times 10^{-2}\ \text{bar}}{0.08314\ \text{bar dm}^3\ \text{mol}^{-1} \times 500\ \text{K}}$$

$$= 4.33 \times 10^{-4}\ \text{mol dm}^{-3}.$$

$$= 4.33 \times 10^{-4}\ \text{mol L}^{-1}.$$

(b) For the reaction $CaCO_3(s) \rightleftarrows CaO(s) + CO_2(g)$

$$\Delta v = 1 - 0 = 1, \quad K_p = 167\ \text{bar at 1073 K}$$

$\therefore$ $$K_c = \frac{K_p}{RT} = \frac{167\ \text{bar}}{0.08314\ \text{bar dm}^3\ \text{mol}^{-1} \times 1073\ \text{K}}$$

$$= 1.872\ \text{mol dm}^{-3}.$$

$$= 1.872\ \text{mol L}^{-1}.$$

S32. For the reaction

(a) $H_2(g) + I_2(g) \rightleftarrows 2HI(g)$

$$K_p(i) = \frac{(pHI)^2}{(pH_2)(pI_2)} \quad \text{... (i)}$$

and for (ii) $2HI(g) \rightleftarrows H_2(g) + I_2(g)$

$$K_p(i) = \frac{(pH_2)(pI_2)}{(pHI)^2} \quad \text{... (ii)}$$

Now, right hand side of equation (ii) may be written as

$$\text{R.H.S.} = \frac{1}{(pHI)^2/(pH_2)(pI_2)} = \frac{1}{K_p(i)}$$

Therefore, K_p (ii) = $1/K_p(i)$.

S33. For the given reaction

$$SO_2(g) + \frac{1}{2}O_2(g) \rightleftarrows SO_3(g)$$

$$K_c = \frac{[SO_3]}{[SO_2][O_2]^{1/2}} = 1.7 \times 10^{12}$$

(a) For reaction $2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g)$

$$K_c(i) = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = 2.89 \times 10^{24}.$$

(a) For reaction $SO_3(g) \rightleftarrows SO_2(g) + \frac{1}{2}O_2(g)$.

$$K_c(ii) = \frac{[SO_2][O_2]^{1/2}}{[SO_3]} = \frac{1}{K_c} = \frac{1}{1.7 \times 10^{12}}$$

$$= 0.588 \times 10^{-12} = 5.88 \times 10^{-13}.$$

S34. Let us suppose that the given reaction is (A) *i.e.*, $N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g)$ ... (A)

(a) $2NH_3(g) \rightleftarrows N_2(g) + 3H_2(g)$ is the reverse reaction. Thus

$$K_p(i) = \frac{1}{K_p(A)} = \frac{1}{41} = 2.439 \times 10^{-2}$$

(b) $\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightleftarrows NH_3(g)$ is half of reaction (A). Thus

$$K_p(ii) = [K_p(A)]^{1/2} = (41)^{1/2} = 6.403$$

(c) $2N_2(g) + 6H_2(g) \rightleftarrows 4NH_3(g)$ is twice of reaction (A), Thus

$$K_p(iii) = [K_p(A)]^2 = (41)^2 = 1681.$$

S35. For the reaction $2\,NO(g) + Br_2(g) \rightleftarrows 2\,NOBr(g)$ at equilibrium

$$[NOBr] = 0.0518 \text{ mol}$$

$$[NO] = [NO]_{initial} - [NO]_{reacted}$$

$$= [NO]_{initial} - [NOBr]_{formed}$$

$$= 0.087 \text{ mol} - 0.0518 \text{ mol} = 0.0352 \text{ mol}$$

$$[Br_2] = [Br_2]_{initial} - [Br_2]_{reacted}$$

$$= [Br]_{initial} - \frac{1}{2}[NOBr]_{formed}$$

$$= 0.0437 \text{ mol} - \frac{1}{2} \times 0.0518 \text{ mol} = 0.0178 \text{ mol}$$

$$K_c = \frac{[NOBr]^2}{[NO]^2\,[Br_2]} = \frac{(0.0518)^2}{(0.0352)^2 \times (0.0178)} = 1.22 \times 10^2.$$

S36. Let us suppose that amount of BrCl used is x.

Therefore, for the reaction:

	$2BrCl(g)$	$\rightleftarrows$	$Br_2(g)$	+	$Cl_2(g)$
Initial mole:	3.30×10^{-3}		0		0
Equilibrium mole:	$(3.30 \times 10^{-3} - x)$		$x/2$		$x/2$

$$K_c = \frac{[Br_2][Cl_2]}{(3.30 \times 10^{-3} - x)^2} = \frac{(x/2)^2}{(3.30 \times 10^{-3} - x)^2}$$

or $$\frac{x/2}{3.30 \times 10^{-3} - x} = \sqrt{K_c} = \sqrt{32} = 5.6568$$

$\therefore$ $$0.5\,x = (3.30 \times 10^{-3}) \times 5.6568 - 5.6568 \times x$$

$$x\,(0.5 + 5.6568) = 3.30 \times 10^{-3} \times 5.6568$$

$\therefore$ $$x = \frac{3.30 \times 10^{-3} \times 5.6568}{6.1568} = 3.0 \times 10^{-3}$$

$$[BrCl] = 3.30 \times 10^{-3} - 3.0 \times 10^{-3}$$

$$= 0.3 \times 10^{-3} = 3.0 \times 10^{-4} \text{ mol L}^{-1}.$$

S37. Initial amount of water = 1 mole

$\therefore$ Initial mass of water = 1 mol $\times$ 18 g mol^{-1} = 18 g

Mass of water that is used up = 40% of the initial mass

$$= \frac{40}{100} \times 18 \text{ g} = 0.4 \times 18 \text{ g} = 7.2 \text{ g}$$

$$\text{Mole of water that is used up} = \frac{\text{Mass of water used up}}{\text{Molar mass of water}}$$

$$= \frac{0.4 \times 18\,g}{18\,g\,mol^{-1}} = 0.4 \text{ mol}$$

According to the balanced chemical equation

Mole of H_2O consumed = Mole of CO consumed = 0.4 mol

Mole of CO_2 formed = Mole of H_2 formed = 0.4 mol

Thus:	$CO(g)$	+	$H_2O(g)$	$\rightleftharpoons$	$CO_2(g)$	+	$H_2(g)$
Initial mole:	1 mol		1 mol		0		0
Equilibrium mole:	(1 – 0.4) mol		(1 – 0.4) mol		0.4 mol		0.4 mol
Equilibrium molarity:	$\frac{0.6\text{ mol}}{10\text{ L}}$		$\frac{0.6\text{ mol}}{10\text{ L}}$		$\frac{0.4\text{ mol}}{10\text{ L}}$		$\frac{0.4\text{ mol}}{10\text{ L}}$

For the given reaction: $$K_c = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(0.4\text{ mol}/10\text{ L}) \times (0.4\text{ mol}/10\text{ L})}{(0.6\text{ mol}/10\text{ L}) \times (0.6\text{ mol}/10\text{ L})} = 0.44$$

S38. Let us suppost that for the reaction $Cl_2 \rightleftharpoons 2Cl$

n_0 = Initial mole of Cl_2

α = Degree of dissociation of Cl_2 at equilibrium

$n_0(1-\alpha)$ = Mole of undissociated Cl_2

$2n_0\alpha$ = Mole of Cl produced

$$\begin{array}{ccc} n_0 & & 0 \\ Cl_2 & \rightleftharpoons & 2\,Cl \\ n_0(1-\alpha) & & 2n_0\alpha \end{array}$$

... initial state

... equilibrium state

Total mole of the mixture = $n_0(1-\alpha) + 2n_0\alpha = n_0(1+\alpha)$

$$\text{Mole fraction of } Cl_2 = \frac{n_0(1-\alpha)}{n_0(1+\alpha)} = \frac{1-\alpha}{1+\alpha} = x(Cl_2)$$

$$\text{Mole fraction of Cl} = \frac{2n_0\alpha}{n_0(1+\alpha)} = \frac{2\alpha}{1+\alpha} = x(Cl)$$

Molar mass of the mixture = $(x \times M)\,Cl_2 + (x \times M)\,Cl$

$$= \frac{1-\alpha}{1+\alpha} \times 71 + \frac{2\alpha}{1+\alpha} \times 35.5$$

$$M(\text{mixture}) = \frac{71}{1+\alpha}$$

According to Graham's law of diffusion, under similar conditions

$$\frac{r(\text{mixture})}{r(\text{Krypton})} = \frac{M(\text{Krypton})}{M(\text{mixture})} = 1.16$$

$\therefore$ $$M(\text{mixture}) = \frac{M(\text{Krypton})}{(1.16)^2} = \frac{84}{(1.16)^2} = 62.42$$

Now, $$M(\text{mixture}) = \frac{71}{1+\alpha} = 62.42$$

or $$71 = 62.42 + 62.42\,\alpha$$

and $$\alpha = 0.1375$$

Fraction of chlorine molecule dissociated = 0.1375.

S39. If total mass of the mixture of CO and CO_2 is 100 g, then

$$\text{Mass of CO} = 90.55\,\text{g}$$

and $$\text{Mass } CO_2 = 100 - 90.55 = 9.45\,\text{g}$$

$\therefore$ $$\text{Number of moles of CO} = 90.55/28 = 3.23$$

$$\text{Number of moles of } CO_2 = 9.45/44 = 0.22$$

$\therefore$ $$p_{CO} = \frac{3.23}{3.23+0.22} \times 1\,\text{atm} = 0.94\text{ atm}$$

$$p_{CO_2} = \frac{0.22}{3.23+0.22} \times 1\,\text{atm} = 0.06\text{ atm}$$

$$K_p = \frac{p_{CO}^2}{p_{CO_2}} = \frac{(0.94)^2}{9.06} = 14.25$$

$$\Delta n_g = 2 - 1 = 1$$

$\therefore$ $$K_p = K_c\,(RT)$$

or $$K_c = \frac{K_p}{RT} = \frac{14.25}{0.0821 \times 1127} = \mathbf{0.154}.$$

S40. $$2N_2(g) + O_2(g) \rightleftharpoons 2N_2O(g)$$

	$2N_2(g)$	$O_2(g)$	$2N_2O(g)$
Initial moles:	0.482	0.933	
Moles at equilibrium:	$(0.482 - x)$	$(0.933 - x/2)$	x

Since, the value of K is very small (2.0×10^{-37}) it indicates that the equilibrium is very much in favour much in favour of reactants and thus the value of x is very very small. Therefore, the approximate concentrations at equilibrium may be writen as

$$[N_2] = \left(\frac{0.482 - x}{10}\right) \approx 0.0482\text{ mol L}^{-1}$$

$$[O_2] = \left(\frac{0.933 - x/2}{10}\right) \approx 0.0933\text{ mol L}^{-1}$$

$$[N_2O] = \left(\frac{x}{10}\right) \quad \text{or} \quad 0.1x \text{ mol L}^{-1}$$

$$K_c = \frac{[N_2O]^2}{[N_2]^2[O_2]} = \frac{(0.1x)^2}{(0.0482)^2(0.0933)}$$

$$= 20 \times 10^{-37}$$

$$x = 6.6 \times 10^{-20}$$

$\therefore$ $[N_2O] = 0.1 \times 6.6 \times 10^{-20} = \mathbf{6.6 \times 10^{-21} \text{ mol L}^{-1}}$

$[N_2] = \mathbf{0.0482 \text{ mol L}^{-1}}$

$[O_2] = \mathbf{0.0933 \text{ mol L}^{-1}}$.

S41. $FeO(s) + CO(g) \rightleftharpoons Fe(s) + CO_2(g)$

	FeO(s)	CO(g)	Fe(s)	$CO_2(g)$
Initial pressures	—	1.4 atm	—	0.80 atm

$$Q_p = \frac{p_{CO_2}}{p_{CO}} = \frac{0.80}{1.4} = 0.571$$

Since $Q_p > K_p$, the net reaction will proceed in the backward direction. Therefore, pressure of CO_2 will decrase and that of CO will increase to attain equilibrium. Hence, if p is the decrease in pressure of CO_2, then increase in pressure of CO = *p*.

$\therefore$ At equilibrium:

$$p_{CO_2} = (0.80 - p) \text{ atm}, \quad p_{CO} = (1.4 + p) \text{ atm}$$

$$K_p = \frac{p_{CO_2}}{p_{CO}}$$

$\therefore$ $0.265 = \frac{0.80 - p}{1.4 + p}$ or $0.265\,(1.4 + p) = 0.80 - p$

or $0.371 + 0.265p = 0.80 - p$

or $1.265p = 0.429$ orr $p = 0.339$ atm

$\therefore$ At equilibrium:

$p_{CO} = 1.4 + 0.339$ atm = **1.739 atm**

$p_{CO_2} = 0.80 - 0.339$ atm = **0.461 atm**.

S42. For the decomposition of x mole of PCl_5, we have

$$PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g)$$

3 mol	0	0		... *Initial*
$3 - x$	x	x		... *Equilibrium*
$3(1-\alpha)$	3α	3α	$\because \alpha = x/3$	

The presence of N_2 will be counted in the calculation of total moles of the gaseous mixture.

$$\text{Equilibrium moles} = \text{Mole of } (PCl_5 + PCl_3 + Cl_2 + N_2)$$

$$= 3(1-\alpha) + 3\alpha + 3\alpha + 1 \text{ mole } N_2$$

$$\therefore \quad n_{total} = 4 + 3\alpha$$

From ideal gas equation $PV = nRt$. we have $n = PV/RT$. Therefore.

$$n_{total} = \frac{PV}{RT}$$

$$4 + 3\alpha = \frac{2.05 \text{ mol} \times 100 \text{ L}}{0.082 \text{ atm L mol}^{-1}\text{K}^{-1} \times 500 \text{ K}} = 5$$

$$\therefore \quad \alpha = (5-4)/3 = 1/3$$

Mole fractions of reactive gases at equilibrium:

$$x(PCl_5) = \frac{3-3\alpha}{5} = \frac{3-3/3}{5} = \frac{2}{5} = 0.4$$

$$x(PCl_3) = \frac{3\alpha}{5} = \frac{3\times 1/3}{5} = \frac{1}{5} = 0.2$$

$$x(Cl_2) = \frac{3\alpha}{5} = \frac{3\times 1/3}{5} = \frac{1}{3} = 0.2$$

Partial pressures of reactive gases at equilibrium:

$$x(PCl_3) = x(PCl_5) \times P = 0.4 \times 2.05 \text{ atm} = 0.82 \text{ atm}$$

$$x(PCl_3) = x(PCl_3) \times P = 0.2 \times 2.05 \text{ atm} = 0.41 \text{ atm}$$

$$p(Cl_2) = x(Cl_2) \times P = 0.2 \times 2.05 \text{ atm} = 0.41 \text{ atm}$$

N_2 is inert, therefore its partial pressure does not appear in the expression of equilibrium constant. Thus,

$$K_p = \frac{p(PCl_3) \times p(Cl_2)}{p(PCl_5)}$$

$$= \frac{(0.41 \text{ atm}) \times (0.41 \text{ atm})}{(0.82 \text{ atm})} = 0.205 \text{ atm.}$$

S43. In the reaction: $COCl_2(g) \rightleftarrows CO(g) + Cl_2(g)$

$$\Delta v = (1+1) - 1 = 1$$

There is increase in the moles as the products are formed. Therefore, according to Le Chatelier's principle an increase of pressure will favour the back direction of the reaction where the number of molecules is less than that in the forward direction. Thus on increasing the pressure:

(a) The equilibrium concentration of CO will decrease.

(b) The partial pressure of $COCl_2$ will increase.

(c) The equilibrium constant (K_p) will not change with change of pressure. Because, this equilibrium constant is related to the partial pressures of the gases in the following form:

$$K_p = \frac{pCO \times pCl_2}{pCOCl_2}$$

The partial pressures of gases are related to their mole fractions as:

$$p(CO) = x(CO) \times P, \qquad p(Cl_2) = x(Cl_2) \times P$$

and $\quad p(COCl_2) = x(COCl_2)\,P$

$$\therefore \quad K_p = \frac{x(CO) \times x(Cl_2)}{x(COCl_2)} \times P = K_x \times P$$

where $\quad K_x = \dfrac{x(CO) \times x(Cl_2)}{x(COCl_2)}$

Now as the pressure is increased the mole fractions of CO and Cl_2 will decrease and that of $COCl_2$ will increase. Thus K_x will decrease. But the product $K_x \times P$ will remain constant when P is increased.

S44. For the given reaction let us suppose that x is the amount of HI dissociated.

Mole of H_2 formed = $x/2$ = Mole of I_2 formed

C_0		0		0	... start
$2\,HI(g)$	$\rightleftharpoons$	$H_2(g)$	+	$I_2(g)$	... equilibrium
$C_0 - x$		$(x/2)$		$(x/2)$	

It is given that $n_0 = 2$ mol and $V = 1$ L. Therefore, $C_0 = 2$ mol L^{-1}.

Equilibrium law gives

$$K_c = \frac{[H_2][I_2]}{[HI]^2} = \frac{[H_2]^2}{[HI]^2} = \frac{[I_2]^2}{[HI]^2}$$

$$\therefore \quad \frac{[H_2]}{[HI]} = (K_c)^{1/2}$$

Thus, $\quad \dfrac{x/2}{C_0 - x} = (K_c)^{1/2}$

$$\frac{x}{2-x} = 2 \times (0.023)^{1/2} = 0.3033$$

On solving for x, we get

$$x = 0.4654 \text{ mol L}-1$$

$\therefore$ $[H_2] = [I_2] = x/2 = 0.2327 \text{ mol L}^{-1}$

and $[HI] = 2 - x = 2 \text{ mol L}^{-1} - 0.4654 \text{ mol L}^{-1}$

$$= 1{,}5346 \text{ mol L}^{-1}$$

$$\alpha = \frac{x}{n_0} = \frac{0.4654}{2} = 0.2327.$$

S45. For the reaction $N_2O_4(g) \rightleftarrows 2NO_2(g)$

Initial mole of $N_2O_4 = n_0 = 1$

Moles of N_2O_4 undissociated $= n_0 - x = 1 - x$

Moles of NO_2 formed $= 2x$

$$\text{Mollarity of } N_2O_4 = [N_2O_4] = \frac{1-x}{V}$$

$$\text{Mollarity of } NO_2 = [NO_2] = \frac{2x}{V}$$

$$\therefore \quad K_c = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(2x/V)^2}{(1-x)/V} = \frac{4x^2}{(1-x)V}$$

or $4x^2 + K_c Vx - K_c V = 0$

It is a quadratic equation. Its solution is given by

$$x = \frac{-K_cV \pm \sqrt{(K_cV)^2 + 16(K_cV)}}{2 \times 4}$$

It is given that $K_c = 4.6 \times 10^{-3}$ and $V = 24.6$ L

$$x = \frac{-4.6 \times 10^{-3} \times 24.6 \pm \sqrt{(4.6 \times 10^{-3} \times 24.6)^2 + 16(4.6 \times 10^{-3} \times 24.6)}}{8}$$

$$= 0.168$$

(a) Degree of dissociation of $N_2O_4 = \alpha = \frac{\text{Moles of } N_2O_4 \text{ dissociated}}{\text{Initial moles of } N_2O_4}$

$$= \frac{0.168 \text{ mol}}{1.0 \text{ mol}} = 0.168$$

(b) Equilibrium concentration of $NO_2 = \frac{2x}{V} = \frac{2 \times 0.168 \text{ mol}}{24.6 \text{ L}}$

$$= 0.0136 \text{ mol L}^{-1}.$$

S46. Initial mole n_0 0

$$N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$$

$n_0 - x$ $2x$

Equilibrium moles	$n_0(1-\alpha)$	$2\,n_0\alpha$
Molarities	$\dfrac{n_0(1-\alpha)}{V}$	$\dfrac{2\,n_0\alpha}{V}$

Equilibrium law gives

$$K_c = \frac{[NO_2]^2}{[N_2O_4]} = \frac{(2\,n_0\alpha/V)^2}{n_0(1-\alpha)/V}$$

$$= \frac{4\alpha^2(n_0/V)}{1-\alpha} = \frac{4\alpha^2 C_0}{1-\alpha}$$

Here $C_0 = n_0/V$ = Initial molarity, and α is the degree of dissociation of N_2O_4.

S47.

Initial mole	n_0		0		0
	$PCl_5(g)$	$\rightleftharpoons$	$PCl_3(g)$	+	$Cl_2(g)$
Equilibrium	$n_0 - x$		x		x
Molarities	$\dfrac{n_0 - x}{V}$		$\dfrac{x}{V}$		$\dfrac{x}{V}$

Equilibrium law gives

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{(x/V)(x/V)}{(n_0 - x)/V} = \frac{x^2}{n_0 - x} \times \frac{1}{V}$$

$$\alpha = \frac{x}{n_0}$$

$\therefore$ $x = n_0\alpha$

$\therefore$ $K_c = \dfrac{(n_0\alpha)^2}{n_0 - n_0\alpha} \cdot \dfrac{1}{V} = \dfrac{\alpha^2}{1-\alpha} \times C_0.$

where $C_0 = n_0/V$ = Initial molarity.

S48. (a) For the reaction $NO(g) + 0.5\,O_2(g) \rightleftharpoons NO_2(g)$

$$\Delta_r G^\ominus = \Delta_f G^\ominus(NO_2) - [\Delta_f G^\ominus(NO) + 0.5\,\Delta_f G^\ominus(O_2)]$$

$$= 52.0\ \text{kJ mol}^{-1} - [87.0 + 0.5 \times 0.0]\ \text{kJ mol}^{-1}$$

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

(b) $$\log K = \frac{\Delta_r G^\ominus}{2.303\,RT}$$

$$= \frac{-(-35.0)\ \text{kJ mol}^{-1}}{2.303 \times 8.314 \times 10^{-3}\ \text{kJ mol}^{-1}\ \text{K}^{-1} \times 298\ \text{K}}$$

$$= 6.134$$

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

S49. (a) $CH_3COOH\ (l) + C_2H_5OH\ (l) \rightleftarrows CH_3COOC_2H_5\ (l) + H_2O\ (l)$

Acetic acid, Ethanol, Ethyl acetate, Water

Reactants: Acetic acid, Ethanol.

Products: Ethylacetate, Water.

(b) The given reaction is an example of liquid phase homogeneous reaction. The concentration ratio of the reaction is given by

$$Q_c = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]}$$

(c) $CH_3COOH\ (l) + C_2H_5OH\ (l) \rightleftarrows CH_3COOC_2H_5\ (l) + H_2O\ (l)$

CH_3COOH	C_2H_5OH	$CH_3COOC_2H_5$	H_2O	
1.00 mol	0.18 mol	0	0	... *Start*
1.00 – 0.171 = 0.829 mol	0.18 – 0.171 = 0.009 mol	0.171 mol	0.171 mol	... *At equilibrium*

$\therefore$ $$K_c = \frac{(0.171\ \text{mol})(0.171\ \text{mol})}{(0.829\ \text{mol})(0.009\ \text{mol})} = 3.92$$

(d) $CH_3COOH\ (l) + C_2H_5OH\ (l) \rightleftarrows CH_3COOC_2H_5\ (l) + H_2O\ (l)$

CH_3COOH	C_2H_5OH	$CH_3COOC_2H_5$	H_2O	
1.00 mol	0.50 mol	0	0	... *Start*
1.00 – 0.214 = 0.786 mol	0.500 – 0.214 = 0.286 mol	0.214 mol	0.214	... *After some time*

$\therefore$ $$Q_c = \frac{0.214 \times 0.214}{0.786 \times 0.286} = 0.204$$

Since $Q_c (0.204) < K_c (3.92)$ for the reaction at this time the equilibrium has not been reached. The reaction will proceed in the forward direction.

(e) Concentrated sulphuric acid absorbs water as soon as it is formed in the reaction. Therefore, according to the Le Chatelier's principle, the reaction proceeds in the forward direction. This causes formation of more products. Moreover, H_2SO_4 acts as catalyst, therefore, formation of ester takes less time.

(f) In the presence of dilute sulphuric acid, the amount of water in the mixture is large. Therefore, equilibrium shifts to the left and the yield of ester is very small.

(g) Since heat of reaction is nearly zero for the given reaction, the equilibrium constant is independent of temperature.

S50. Mole of I_2 in one litre (*i.e.*, 1000 mL) of a saturated solution = 1.1×10^{-3} mol

$\therefore$ Mole of I_2 in 100 mL of saturated solution $= 1.1 \times 10^{-3}\ \text{mol} \times \frac{100\ \text{mL}}{1000\ \text{mL}}$

$= 1.1 \times 10^{-4}$ mol

Now, molar mass of $I_2 = 1 \times I = 2 \times 127 = 254$ g/mol

(a) Mass of iodine in 100 mL of saturated solution

$= (\text{mole} \times \text{molar mass}) \text{ of } I_2$

$= 1.1 \times 10^{-4} \text{ mol } 254 \text{ g/mol}$

$= 279.4 \times 10^{-4} \text{ g} = 0.02794 \text{ g} = 0.028 \text{ g}.$

(b) Initial mass of iodine = 200 mg = $\frac{200}{1000}$ = 0.200 g

Mass of undissolved iodine = Initial mass – Disdsolved mass

= 0.200 g – 0.028 g = 0.172 g.

(c) Total volume of solution = 100 mL + 150 mL = 250 mL

(i) Mass of iodine in 100 mL saturated solution = 0.028 g

Mass of iodine in 250 mL saturated solution

$$= 0.028 \text{ g} \times \frac{250 \text{ mL}}{100 \text{ mL}} = 0.07 \text{ g}$$

(ii) Mass of undissolved iodine = 0.200 g – 0.07 g = 0.130 g

(iii) The concentration of the solution (mol/L) will be same as in the original saturated solution.

$$\text{Molarity} = \frac{\text{Mole of solute}}{\text{Volume of solution}} = \frac{0.07 \text{ g}}{254 \text{ g/mol}} \times \frac{1}{0.250 \text{ L}}$$

$= 1.1 \times 10^{-3}$ mol/L..

S51. It is given that for $N_2O_4(g) \rightleftarrows 2NO_2(g)$.

$$K_p = 80 \text{ atm} = \frac{(pNO_2)^2}{pN_2O_4}$$

(a) $$Q_p = \frac{(4 \times 10^{-3} \text{ atm})^2}{4 \times 10^{-3} \text{ atm}} = 0.004 \text{ atm} < K_p$$

Since $Q_p < K_p$, the reaction is not at equilibrium. The amount of NO_2 is less than the amount of N_2O_4. Therefore, the reaction will proceed in the forward direction so that N_2O_4 is changed into NO_2 and the equilibrium is attained, and $Q_p = K_p$.

(b) $$Q_p = \frac{(0.4 \text{ atm})^2}{2 \times 10^{-3} \text{ atm}} = 80 \text{ atm} = K_p$$

Since $Q_p = K_p$, the system represents the state of equilibrium.

(c) $$Q_p = \frac{(1 \text{ atm})^2}{(0.002) \text{ atm}} = 500 \text{ atm} > K_p$$

Since $Q_p > K_p$, the system does not represent the state of equilibrium. The amount of NO_2 is much greater than the amount of N_2O_4. Therefore, the reaction will proceed in the back direction and NO_2 will be changed to N_2O_4 so that equilibrium is attained and $Q_p = K_p$.

(d) $$Q_p = \frac{(2 \times 10^{-3} \text{ atm})^2}{0} = \infty \gg K_p$$

The amount of N_2O_4 is zero and $Q_p \gg K_p$. Therefore, NO_2 will be consumed to form N_2O_4 so that equilibrium is reached and $Q_p = K_p$.

S52. For the reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ ... (i)

$$K_p = \frac{p^2(\text{HI})}{p(H_2) \times p(I_2)}.$$

It is given that each mixture contains solid I_2. As I_2 vapour is consumed in the reaction (i) solid I_2 sublimes $I_2(s) \rightleftharpoons I_2(g)$ and more I_2 vapour is produced. There is no change in the amount of $I_2(g)$ and its partial vapour pressure remains constant as long as solid I_2 is present in the container. Thus, $p(I_2)$ is constant in each container. Let us suppose that $p(I_2) = x$ = constant. Therefore,

$$\frac{p^2(\text{HI})}{p(H_2) \times x} = Q_p$$

(a) $p(H_2) = 0.10$ atm, $p(\text{HI}) = 0.80$ atm, $p(I_2) = x$ = constant

$\therefore$ Partial pressure quotient = $Q_p = \frac{p^2(\text{HI})}{p(H_2) \times x} = \frac{(0.80)^2}{0.1 \times x} = \frac{6.4}{x}$

or $$Q_p \times x = Q_p' = 6.4$$

$$Q_p' > 0.35.$$

Here, we find that $Q_p'(6.4) > K_p(0.35)$. Therefore, the reaction will proceed in the back direction so that HI decomposes to form H_2 and I_2.

(b) $p(H_2) = 0.55$ atm, $p(\text{HI}) = 0.44$ atm, $p(I_2) = x$ = constant

$\therefore$ $$Q_p = \frac{p^2(\text{HI})}{p(H_2) \times p(I_2)} = \frac{(0.44)^2}{0.55 \times x} = \frac{0.35}{x}$$

or $$Q_p \times x = 0.35 = Q_p' = K_p$$

$$Q_p' = 0.35.$$

Since $Q_p' = K_p$, therefore, the reaction is in the ***state of equilibrium***.

(c) $p(H_2) = 2.5$ atm, $p(\text{HI}) = 0.15$ atm, $p(I_2) = x$ = constant

$\therefore$ $$Q_p = \frac{p^2(\text{HI})}{p(H_2) \times p(I_2)} = \frac{(0.15)^2}{2.5 \times x} = \frac{0.009}{x}$$

$\therefore$ $$Q_p \times x = 0.009 = Q_p' < 0.35$$

$$Q_p' < 0.35.$$

Since $Q_p' < K_p$, the reaction will proceed in the ***forward direction*** so that more HI is formed by the consumption of H_2 and I_2.