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TYPES OF CHEMICAL REACTIONS

+ Reversible reaction: Reaction in which reactant changes to product but product can also give back reactants.

$$A + B \longrightarrow AB \qquad AB \longrightarrow A + B$$
$$A + B \rightleftharpoons AB$$

It ~~also~~ has two reactions happening at the same time. Forward direction & Backward direction.

+ Rate of reaction: Defined as the change in concentration of reactant or product per unit time. There is drastic change in concentration at beginning which slowly decreases with time. So rate is high at start of reaction but is low at ends.

$$A + B \rightleftharpoons C + D$$

$A + B \xrightarrow{i} C + D$ ← Now there's no backward reaction because A and B have high concentration and C and D are just being formed.

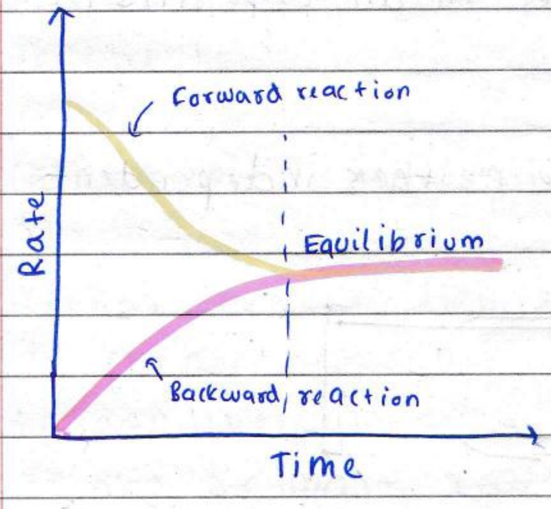
$C + D \xrightarrow{ii} A + B$ ← Now as C and D are produced, backward reaction also becomes possible.

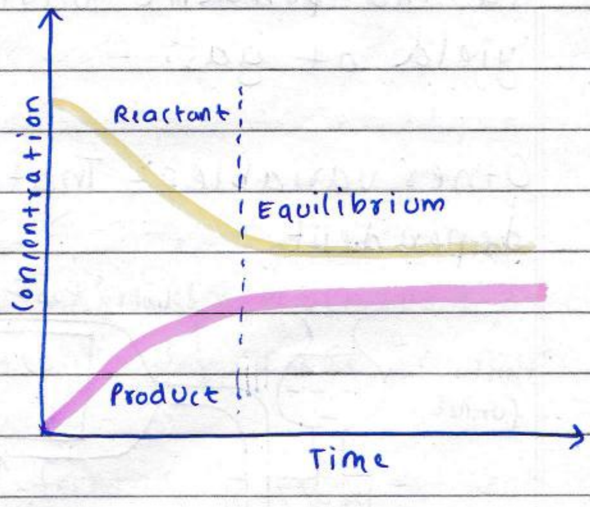
Rate of reaction i goes on decreasing while that of ii goes on increasing with time.
When Rate i = Rate ii, this is called equilibrium.
For first few seconds rate of forward reaction is greater than rate of backward reaction.

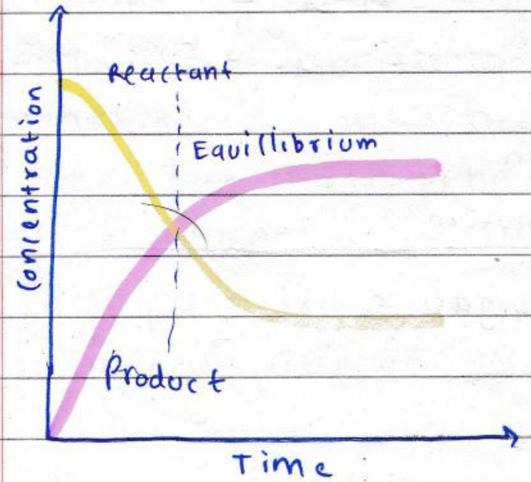
Equilibrium is the condition of a reversible reaction at which rate of forward reaction becomes equal to rate of backward reaction and concentrations of both reactants and products donot change. It is ultimately dynamic as reactants change to products and products change to reactants.

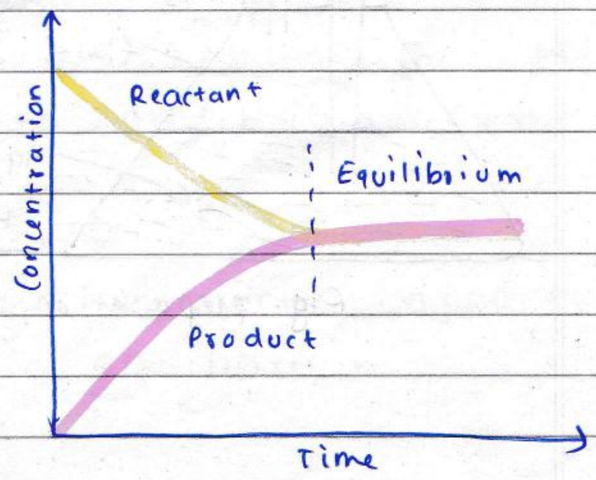
Characteristics of equilibrium

- It is dynamic in nature.
- Rate of forward reaction is equal to rate of backward reaction.
- Concentration of reactants and products don't change with time.
- Requires closed system.









An equilibrium can be disturbed. After this, the direction the equilibrium follows is ~~poo~~ called position of equilibrium / shifting of equilibrium. It refers to the relative amounts of reactants and products in an equilibrium.

Rate of reaction is directly proportional to concentration of reactant and temperature.

Independent variable: Variables that we determine and change in an experiment. Eg: mass of reactant, concentration of reactant, temperature of reaction etc.

Dependent variable: Variables that depends on the independent variables in an experiment. Eg: yield of gas.

Other variables: That are neither independent or dependent.

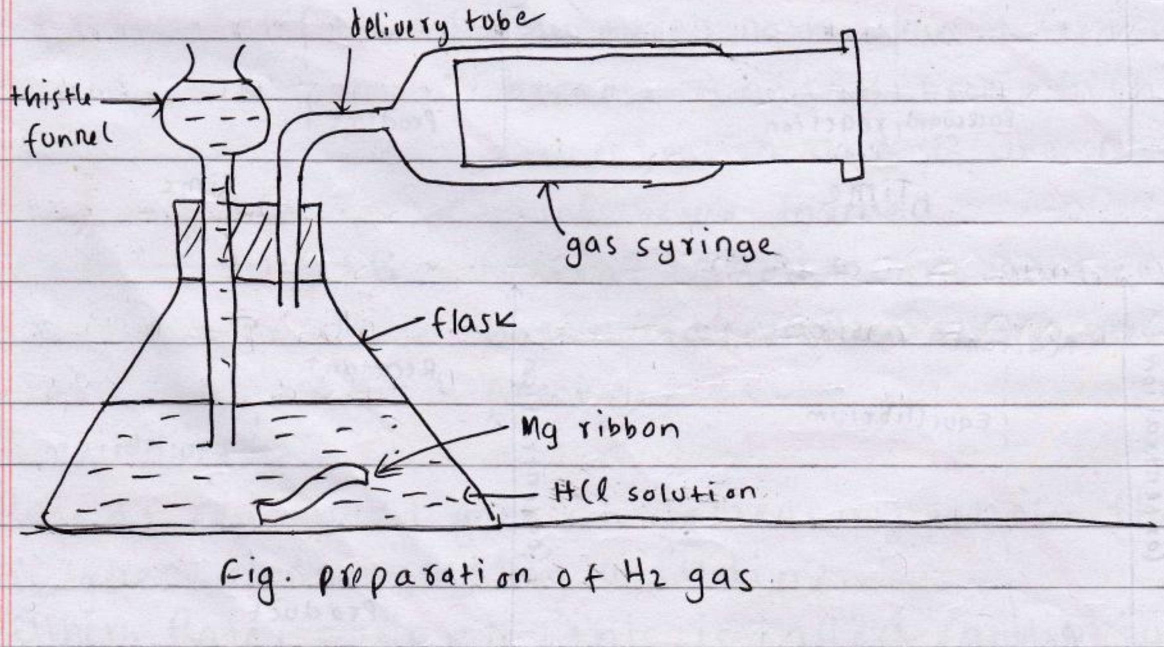


Fig. preparation of H_2 gas

Position of equilibrium

Refers to the relative amounts of products and reactants present in an equilibrium mixture.

$A + B \rightleftharpoons C + D$

- If a dynamic equilibrium is disturbed by changing the conditions, position of equilibrium moves to counteract the change.

$A + 2B \rightleftharpoons C + D$

Let concentration of A be $[A]_{eq}$ at equilibrium.

If $[A]_{eq} + [B]_{eq} > [C]_{eq} + [D]_{eq}$, position of equilibrium is towards left / backwards / ~~to~~ reactants.

Le Chatelier's principle

- French Chemist Henri Le Chatelier observed effect of changes in concentration, pressure, temp. etc. to equilibrium.
- If one or more ~~reaction~~ factors that affect a dynamic equilibrium is changed, position of equilibrium shifts to nullify this change. Then the reaction again returns to equilibrium after some time but with changes.
- These changes are effects.
- This principle says that the system opposes what we do, to stay in equilibrium.

When concentration of reactants in increased.

- system is no longer in equilibrium.
- position of equilibrium shifts to the right.

$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$

ethanoic acid ethanol ethyl ethanoate water.

If more ethanol is added, system tends to decrease the concentration by shifting equilibrium to product side, thus causing reaction between ethanoic acid and excess ethanol decreasing ethanol's concentration, again resolting in equilibrium where concentrations are disimilar.

$A + B \rightleftharpoons C + D$

If concentration of D is increased, rate of backward reaction is enhanced, thus position of equilibrium shifts to left. This consumes some amount of C and D to form more A and B. Thus system arrives in equilibrium

$Ce^{4+} + Fe^{2+} \rightleftharpoons Ce^{3+} + Fe^{3+}$

If Fe^{2+} ions are increased, rate of forward reaction is enhanced to produce more products and less reactants. Thus equilibrium shifts to product side.

Effect of pressure on position of equilibrium

- The pressure of a gas is caused by molecules hitting on walls of container. So, at constant temperature, more gas molecules there are ie. more collision, more volume higher will be pressure.
- When pressure increases, equilibrium shifts towards side of reaction with fewer moles of gas and vice versa.

$A + 2B \rightleftharpoons 2C$

3 moles   2 moles

- If position of equilibrium is shifted to right by us, an immediate effect is that we get more product, whis has lesser no. of mols, which means lesser ~~pres~~ volume and thus lesser pressure. ~~So, when~~
- However this rule is only applicable to gases.

Effect of temperature

$H_2 + CO_2 \rightleftharpoons H_2O + CO \quad \Delta H_r = +41.2\ kJ\ mol^{-1}$

- If we increase temp. system tends to decrease it, thus system shifts position of equilibrium in direction of endothermic reaction / forward reaction.

Catalysts have no impact as they enhance rate of forward and rate of backward reactions equally.

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EQUILIBRIUM EXPRESSION AND EQUILIBRIUM CONSTANT (K_c and K_p):

For a reaction:

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

Equilibrium not changed by pressure as 1 moles = 1 mole

Note: We mustn't say that temperature and/or pressure of reactants is changed. It must be said that temperature and/or pressure of system is changed.

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92\ kJ\ mol^{-1}$

When temperature of system is changed system is no longer in equilibrium. According to Le Chatelier principle system opposes what we do.

Equilibrium constants (K_c/K_p):

$$aA + bB \rightleftharpoons cC + dD$$

Equilibrium concentrations are represented by [ ] in $mol\ dm^{-3}$. Thus, equilibrium concentration of A = [A] or [A] eqm.

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \frac{\text{product}}{\text{reactant}}$$ This is equilibrium expression

∴ Equilibrium constant (K_c) is defined as the ratio of product of molar concentrations of products to product of molar concentrations of reactants with each concentration term having power equal to its stoichiometric coefficient, given by a balanced chemical eqⁿ.

$N_2 + 3H_2 \rightleftharpoons 2NH_3$ # $2SO_2 + O_2 \rightleftharpoons 2SO_3$

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} \quad K_c = \frac{[SO_3]^2}{[O_2][SO_2]^2}$$

When temperature of system is lowered, system is no longer in equilibrium and opposes what we do. It then increases the temperature by shifting reaction in forward direction i.e towards $2NH_3$ (exothermic) which forms more NH_3.

When pressure of system is ~~mad~~ increased, system is no longer in equilibrium and opposes what we do. It then decreases the pressure by shifting reaction towards direction with lesser number of moles, i.e towards $2NH_3$, which forms more NH_3

Equilibrium ~~co~~ expression:
Relationship between the equilibrium constant (K_c) ~~and~~ equilibrium molar concentrations of reactants and products and their stoichiometric coefficients.

Equilibrium expression in terms of K_p.
Relationship between the equilibrium constant (K_p), equilibrium ~~molar concentr.~~ partial pressures of reactants and products and their stoichiometric coefficients.

$$K_p = \frac{p_C^c \cdot p_D^d}{p_A^a \cdot p_B^b}$$, where p_C is equilibrium partial pressure of C.

The reason for gas pressure is because gas molecules have negligible intermolecular forces because we assume temp. is enough such that KE of molecules can overcome intermolecular forces of attraction. The perfectly elastic collision of molecules with walls of container and each other.

EQUILIBRIUM EXPRESSION...

- Gas exerts pressure due to the bombardment of gaseous molecules on the walls of the container. The pressure exerted by gas depends on temperature at constant volume. Directly proportional as if temp. increases, molecules strike walls with a higher velocity. This gives total pressure.
- Partial pressure is the amount of pressure exerted by the individual gases in a gaseous mixture.
- Total pressure is sum of all partial pressures.

In a mixture with A, B, C, D gases, let's suppose the no. of moles are n_1, n_2, n_3, n_4

Total pressure = 100 Pa.

Mole fraction of A (n_A) = $\dfrac{n_1}{n_1+n_2+n_3+n_4\ (n)}$

∴ Partial pressure of A (p_A) = $\dfrac{n_1}{n} \times P_t$

$= n_A \times 100$

END OF CHAPTER QUESTIONS

1. The reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ reaches dynamic equilibrium in a closed vessel. The forward reaction is exothermic. The reaction is catalyzed by V_2O_5.

a) Explain dynamic equilibrium.

→ Dynamic equilibrium is a condition of reversible reaction in which the rate of forward reaction is equal to the rate of backward reaction and the concentrations of reactants and products remain constant. In a dynamic equilibrium, reactants constantly change into products and vice versa.

b) What will happen to position of equilibrium if:

i) Some SO_3 is removed from the vessel.

→ As concentration of SO_3 is lowered, system is no longer in equilibrium and tries to oppose what we do by increasing concentration of SO_3 by enhancing rate of forward reaction. Thus, position of equilibrium shifts to right.

ii) the vessel pressure is lowered.

→ The system is no longer in equilibrium and opposes what we do by increasing pressure, by enhancing rate of backward reaction (to more number of moles). Thus, position of equilibrium shifts to left.

iii) more V_2O_5 is added?

→ The position of equilibrium is unchanged as catalyst enhances rate of both forward and backward reactions.

iv) the temp. of the vessel is increased.

→ The system is no longer in equilibrium, and opposes what we do by lowering temperature, which is done by enhancing rate of backward reaction which is endothermic. Thus, position of equilibrium shifts to the left.

c) State Le chatelier's principle.

→ Le chatelier's principle states that, "if one of the factors changing a dynamic equilibrium like concentration, pressure or ~~p~~ temperature is changed, then the system opposes what we do inorder to nullify that effect."

d) Use Le chatelier's principle to explain what will happen to position of equilibrium in reaction

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

when concentration of hydrogen is increased.

→ When conc. of H_2 is increased, then the system is no longer in equilibrium. The system then opposes what we do by attempting to ~~increase~~ decrease conc. of H_2. This is done by shifting position of equilibrium to the ~~left~~ right, enhancing rate of ~~backward~~ forward reaction, thus producing ~~more~~ less H_2.

2. Hydrogen, iodine and hydrogen iodide are in equilibrium.

$H_2 + I_2 \rightleftharpoons 2HI(g) \quad \Delta H_r = -J6 \text{ kJ mol}^{-1}$

The partial pressurer of each gas is shown.

Gas	Partial Pressure / Pa
H_2	2.33×10^6
I_2	0.925×10^6
HI	10.2×10^6

a) Explain the meaning of partial pressure.

→ Partial pressure is the pressure exerted by a particular gas to the walls of the container it is kept in, in a gaseous mixture.

b) Calculate total pressure of all three gases.

→ $P_{total} = P_A + P_B + P_C$

$= (2.33 \times 10^6) + (0.925 \times 10^6) + (10.2 \times 10^6)$

$= 1.34 \times 10^7$ Pa //

c) Write an equilibrium expression for this reaction in terms of partial pressures.

→ $K_p = \dfrac{p_{HI}^{\,2}}{p_{H_2} \cdot p_{I_2}}$

d) Calculate a value of Kp for this reaction.

→ $K_p = \dfrac{(10.2 \times 10^6 \text{ Pa})^2}{(2.33 \times 10^6 \text{ Pa})(0.925 \times 10^6 \text{ Pa})}$

$= \dfrac{1.0404 \times 10^{14}}{2.15525 \times 10^{12}}$

$= 48.27$ //

e) Use Le Chatelier's principle to explain what happens to position of equilibrium when:

i) the temp. is increased.

→ The system is no longer in equilibrium and opposes what we do by decreasing the temp. This is done by enhancing rate of backward reaction which is endothermic. Thus, position of equilibrium shifts to the left.

ii) Some iodine is removed,

→ As conc. of I_2 decreases, system is no longer in equilibrium and opposes what we do by increasing conc. of I_2 by enhancing rate of backward reaction. Thus, position of equilibrium shifts to the left.

3. The equilibrium between three substances is shown.

$A(g) + B(g) \rightleftharpoons C(g)$

Initially there was 0.1 mol A, 0.2 mol B in mixture which reacted to form a equilibrium mixture with 0.04 mol C. The total volume of mixture was 2.00 dm^3.

a) Calculate no. of moles of A and B at equilibrium.

→ ~~Volume of C = 0.04 × 24 = 0.96 dm^3~~

~~Volume of A, B = 2 - 0.96 = 1.04 dm^3~~

~~$24(n_A + n_B) = 1.04$~~

~~$\therefore n_A + n_B = 13/300$~~

0.04 mol A and 0.04 mol B reacted to form 0.04 mol C [1 : 1 : 1]

∴ At equilibrium,

∴ $n_A = 0.1 - 0.04 = 0.06$ mol

∴ $n_B = 0.2 - 0.04 = 0.16$ mol

b) Calculate the concentrations of A, B and C at equilibrium.

~~$0.06n + 0.16n + 0.04n = 2$~~

~~$0.26\,n = 2$~~

~~$\therefore n = 100/13$~~

$n = cV$, $c = n/V$

$\therefore C_A = n_A/V = 0.06/2 = 0.03 \text{ mol dm}^{-3}$

$\therefore C_B = n_B/V = 0.16/2 = 0.08 \text{ mol dm}^{-3}$

$\therefore C_C = n_C/V = 0.04/2 = 0.02 \text{ mol dm}^{-3}$

ci) Write the equilibrium expression for Kc.

$$\therefore K_c = \frac{[C]}{[A][B]}$$

ii) Calculate value of Kc and give units.

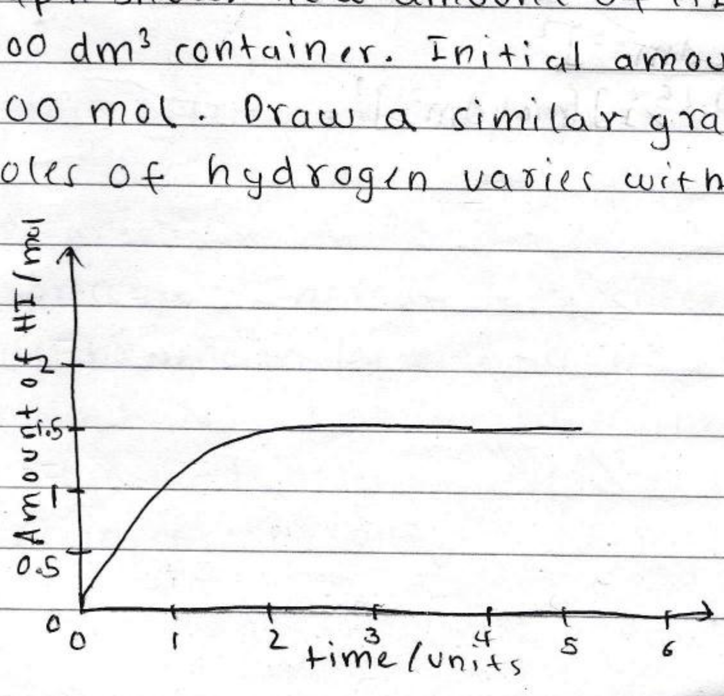
$$\therefore K_c = \frac{[0.02 \text{ mol dm}^{-3}]}{[0.03 \times 0.08](\text{mol dm}^{-3})^2}$$

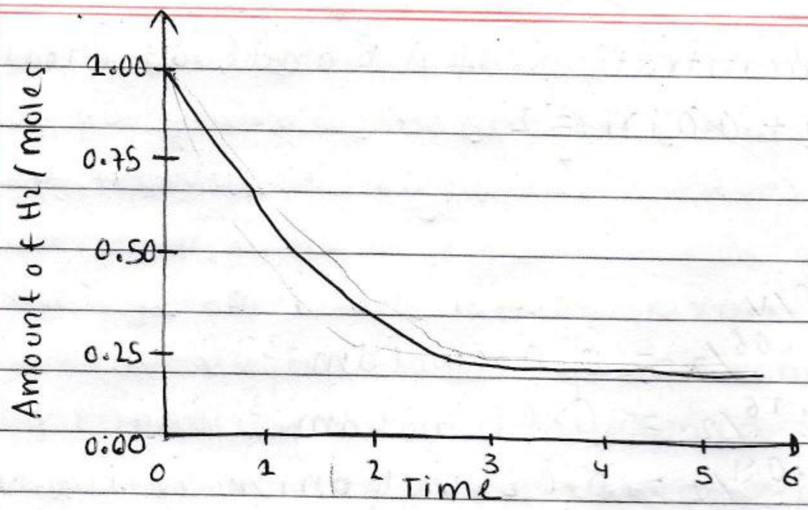
$$= 8.33 \text{ mol}^{-1}\text{dm}^{3}$$

4. Gaseous hydrogen and iodine react to form HI.

$$H_2 + I_2 \rightleftharpoons 2HI$$

a) Graph shows how amount of HI varies with time in a 1.00 dm^3 container. Initial amounts of H_2 and I_2 were 1.00 mol. Draw a similar graph to show how no. of moles of hydrogen varies with time.





$n(HI) = 1.5$ moles [At equilibrium]

$\therefore n(H_2) = 1.5/2 = 0.75$ moles [At equilibrium]

$\therefore n(H_2) = 1 - 0.75 = 0.25$ moles [Remaining]

b) Calculate no. of moles of iodine present at equilibrium.

$n(I_2) = n(H_2)$ [1 : 1]

$\therefore n(I_2) = 0.25$ moles.

ci) Write equilibrium expression for Kc for reaction.

$\rightarrow$ $Kc = \dfrac{[HI]^2}{[H_2][I_2]}$

ii) Calculate value of Kc with units.

$\rightarrow$ $Kc = \dfrac{[1.5/1 \text{ mol } \cancel{dm^{-3}}]^2}{[0.25/1][0.25/1][\text{mol } \cancel{dm^{-3}}]^2}$

$= \dfrac{1.5^2}{0.0625}$

$= \cancel{24}$ 36

5a) Describe three characters of chemical equilibrium.
→ It is dynamic in nature.
→ Rate of forward reaction is equal to rate of backward reaction.
→ Concentration of reactants and products don't change with time.

b) When 1 mol N_2O_4 is allowed to come to equilibrium with NO_2 gas under standard conditions, only 20% N_2O_4 is converted to NO_2.

$$N_2O_4 \rightleftharpoons 2NO_2 \quad \Delta H_r = +58\ kJ\ mol^{-1}$$

i) Give equilibrium expression for reaction.
→ $K_c = \frac{[NO_2]^2}{[N_2O_4]}$

ii) Calculate value of Kc ~~with~~ assuming volume is $1.00\ dm^3$.
→ $K_c = \frac{[2 \times 0.2 \times 1/1]^2}{[0.8 \times 1/1]}$

$= \frac{0.16}{0.8}\ mol\ dm^{-3}$

$= 2\ mol\ dm^{-3}$

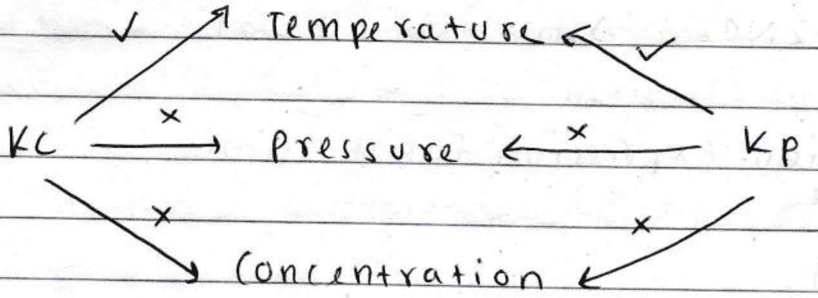
c) Explain effect on Kc of an increase in:

i) pressure
→ As pressure is increased, the system shifts position of equilibrium to left to decrease pressure. This decreases conc. of NO_2 but increases that of N_2O_4. This decrease and increase balance each other, thus Kc remains constant.

ii) temperature

→ As temperature rises, system shifts position of equilibrium to right to endothermic side. As temp. is exponentially proportional to rate of reaction, the rate of change of concentrations can't balance each other, and thus K_c is increased.

EFFECT OF TEMPERATURE, PRESSSURE AND CONCENTRATION ON K_c and K_p



If concentration of reactants is increased, system shifts position of equilibrium to product side to produce more product. But the value by which product increases is same as value by which reactant increases as all extra reactants don't react. This ratio causes K_c to remain constant.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

1 mole   2 mole   3 mole   — R_1

1 mole N_2 added, 0.75 mole reacted, 0.25 mole remains.

1.25 mole   2 mole   ~~3.325 mole~~ 4.5 moles   [1:2]   — R_2

The ratio R_1 and R_2 are made constant thus value of equilibrium constant is unchanged.

What is the effect of change in conc. press. & temp in:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$. $\Delta H_r = -ve$

If pressure of system is increased, total pressure also increases, and thus partial pressure also increases in same ratio making Kp constant.

When the temperature of system is increased, system no longer remains in equilibrium. Pos. of equilibrium shifts to the endothermic side. This cause conc. of reactants to decrease but products to increase. This thus causes Kc to increase as well as $K_C = \frac{\text{Products}\uparrow}{\text{reactants}\downarrow}$

$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

$K_C = [CO_2]$

This is because concentration of solids are always the same. They will only change if the molecular separation is changed.

0.30 mol of dinitrogen tetraoxide is allowed to come to equilibrium with nitrogen dioxide. The amount of nitrogen dioxide at equilibrium is 0.28 mol. Calculate amount of dinitrogen tetraoxide at equilibrium, Kc, given that the volume of the containing vessel is 10.0 dm³

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

→ $n(N_2O_4) = 0.28/2 = 0.14$ mol ∴ $n = 0.30 - 0.14 = 0.16$ mol

→ $K_C = \frac{[0.16/10]^2}{[0.28/10]^2} = \cancel{0.33 \text{ mol dm}^{-3}}$, $\frac{1}{\cancel{0.33}} = 3.06\ \cancel{\text{mol}}\ \text{dm}^{-3}$

→ $K_C = \frac{[0.28/10]^2}{0.16/10} = 0.049 \text{ mol dm}^{-3}$

In Haber's process ammonia is produced as:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H_r = -92\ kJ\ mol^{-1}$

The reaction is carried out initially by taking 0.5 mol nitrogen and 1.5 mol hydrogen. At equilibrium, 0.6 mol of NH_3 is formed. given that the volume of $5\ dm^3$. Calculate equilibrium concentrations of N_2, H_2 and NH_3 and find K_c. What is effect of increasing temp. on K_c. What are conditions for max. yield of NH_3.

~~$[N_2]_{eqm} = {}^{0.6/2}/5 = {}^{0.3}/5 = 0.06\ mol/dm^{-3}$~~

~~$[H_2]_{eqm} = {}^{3\times0.3}/5 = \cancel{0.12\ mol/dm^{-3}}\ 0.18\ mol\ dm^{-3}$~~

~~$[NH_3]_{eqm} = {}^{0.6}/5 = 0.12\ mol\ dm^{-3}$~~

~~$K_c = \frac{[0.12]^2}{[0.06][0.18]^3} = 41.2\ mol^{-2}\ dm^6$~~

$[N_2]_{eqm} = \frac{0.5 - 0.6/2}{5} = \frac{0.2}{5} = 0.04\ mol\ dm^{-3}$

$[H_2]_{eqm} = \frac{1.5 - 3\times0.3}{5} = \frac{0.6}{5} = 0.12\ mol\ dm^{-3}$

$[NH_3]_{eqm} = \frac{0.6}{5} = 0.12\ mol\ dm^{-3}$

$\therefore K_c = \frac{[0.12]^2}{[0.04][0.12]^3} = 208.3\ mol^{-2}\ dm^6$

Conditions for max. yield of ammonia are low temperature and high pressure.

Equilibrium is shifted towards the left.

END OF CHAPTER QUESTIONS...

6. $CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$

9.2 g ethanol mixed with 12.0 g of ethanoic acid in 250 cm^3 is left to equilibriate. At eqm. 70% reactants turn to products.

a) What is concentration of each reactant at start?

$n = \frac{m}{M} = \frac{9.2}{46} = 0.2$ moles $\quad n = \frac{m}{M} = \frac{12}{60} = 0.2$ moles.

$\therefore [CH_3COOH] = 0.2/0.25 = 0.8 \text{ mol dm}^{-3}$

$\therefore [C_2H_5OH] = 0.2/0.25 = 0.8 \text{ mol dm}^{-3}$

b) What is concentration of each reactant at equilibrium.

70% of 0.2 = $7/10 \times 0.2 = 0.14$ moles.

Remaining moles = 0.2 - 0.14 = 0.06 moles.

$\therefore [CH_3COOH]_{eqm} = 0.06/0.25 = 0.24 \text{ mol dm}^{-3}$

$\therefore [C_2H_5OH]_{eqm} = 0.06/0.25 = 0.24 \text{ mol dm}^{-3}$

c) What is concentration of each product at equilibrium.

$n[CH_3COOC_2H_5] = 0.14$ moles $\quad n[H_2O] = 0.14$ moles.

$\therefore [CH_3COOC_2H_5]_{eqm} = 0.14/0.25 = 0.56 \text{ mol dm}^{-3}$

$\therefore [H_2O]_{eqm} = 0.14/0.25 = 0.56 \text{ mol dm}^{-3}$

d i) Write equilibrium expression for this reaction.

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

ii) Calculate value of Kc for this reaction.

$$K_c = \frac{[0.56][0.56]}{[0.24][0.24]} = 5.44$$

iii) Explain why K_c has no units.
→ K_c has no units as molar concentrations are ~~not~~ raised to ~~a~~ power of 1 and there are two reactants and two products.

7a) Hydrogen chloride and ammonia, both ionize in water.
$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$
$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

i) State the name of the ion H_3O^+
→ Hydronium ion.

ii) Identify acid and base on LHS of both eqns.
→ HCl is acid while NH_3 is base.

iii) Explain why water is amphoteric.
→ As in first eqn water acts as a base to give H_3O^+ and in second eqn it acts as an acid to give OH^-, it is amphoteric.

~~b) When dissolved in an~~

8. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H_r = -92 \text{ kJ mol}^{-1}$
120 mol of H_2 is mixed with 40 mol of N_2 then its pressurized. The mixture of gases is allowed to equilibriar. Total volume is 1.00 dm^3. 20% of reactants become NH_3.

a) How many moles of N_2 and H_2 remain at equilibrium.
20% of 40 = 8 moles. 20% of 120 = 24 moles.
→ $n(N_2)$ eqm = 40 − 8 = 32 moles.
→ $n(H_2)$ eqm = 120 − 24 = 96 moles.

b) How many moles of NH_3 are formed.
→ $n(NH_3)_{eqm} = 2 \times 8 = 16$ moles.

c) Write an eqm. expression for Kc.

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

d) Calculate value of Kc.

$$K_c = \frac{[16]^2}{[32][96]^3} = 9.04 \times 10^{-6}\ mol^{-2}\,dm^{6}$$

e) What will happen to Kc when pressure is raised.
→ When pressure is raised, the system is no longer in equilibrium. System shifts pos. of eqm to the right. Thus number of moles of NH_3 increases, increasing molar concentration of NH_3. But as volume is reduced to raise pressure, the molar concentrations of N_2 and H_2 also increases so as to balance out. Thus Kc remains constant.

f) What will happen to Kc when temp. is raised.
→ When temp. is raised, system is no longer in equilibrium. System shifts pos. of eqm to the left. Thus number of moles of N_2 and H_2 increases but that of NH_3 decreases. Thus molar concentrations of N_2 and H_2 increases while that of NH_3 decreases. Thus value of Kc also decreases.

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9. Ethanol can be manufactured by reacting C_2H_4 with steam.
$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$

a) Write eqm. exp. in terms of Kp.
→ $K_p = \frac{p_{C_2H_5OH}}{p_{C_2H_4} \cdot p_{H_2O}}$

b) State units for Kp in this reaction.
→ Unit is Pa^{-1}.

c) Reaction is at eqm at 290°C and 7×10^6 Pa pressure. Under these conditions the partial pressure of ethene is 1.5×10^6 Pa and partial pressure of steam is 4.2×10^6 Pa.

i) Calculate partial pressure of ethanol.
→ $p_{ethanol} = (7 \times 10^6) - (1.5 \times 10^6) - (4.2 \times 10^6) = 1.3 \times 10^6$ Pa

ii) Calculate value of Kp under these conditions.
→ $K_p = \frac{1.3 \times 10^6}{1.5 \times 10^6 \times 4.2 \times 10^6} = 0.21\ Pa^{-1}$

d) The reaction is carried out in a closed system. Explain closed system.
→ A system where transfer of energy can take place but transfer of matter can't is closed system.

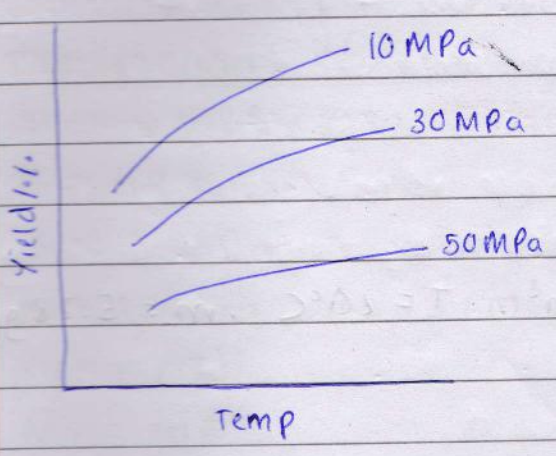
e) What will happen to pos. of eqm ~~in~~ when pressure is increased
→ The system is no longer in equilibrium. To decrease the pressure, the position of eqm shifts to side with lesser number of moles. Thus pos. of eqm shifts to the right.

f) Use the given info to explain sign of enhalpy change for this reaction.

Temp/°C	% of ethene converted
260	40
290	38
320	36

As temp increases, no. of moles of C_2H_5OH decreases. Thus system shifts eqm to the left. Therefore this forward reaction is exothermic.

#



Effect of temp. and pressure on yield of product in eqm. Deduce wether forward reaction involves and increase or decrease in number of moles.

As pressure is increased, yield% decreases which means eqm. shifts to backward reaction. This means backward reaction has lesser number of moles. Thus the products of forward reaction produces more number of moles. Therefore, forward reaction involves an increase in number of moles.

As temp is increased, yield% increases which means pos. of eqm. shifts to the right. Thus forward reaction reduces temp. and is endothermic.

NO_3^{-1}

$n + 3x - 2 = -1$

$n - 6 = -1 \quad n = +5$

$2x + 2 - 6 = 0$

$2x =$

EQUILIBRIUM...

\# 0.218 mol of HI was heated in a flask of V dm^3, where eqm. was established at 700 K.

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

The eqm. mixture was found to have 0.023 mol H_2.

i) Calculate no. of moles of iodine and hydrogen iodide in eqm mixture.

$n(I_2) = 0.023$ mol

$n(HI) = 0.218 - 2 \times 0.023 = 0.172$ mol

ii) Write an expression for Kc.

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

iii) Why is volume not needed while calculating Kc.

→ Volume is not needed as in the expression for Kc volume gets cancelled out. $V.V/V^2 = 1$ //

iv) Calculate Kc.

$$Kc = \frac{[0.023][0.023]}{[0.172]^2} = 0.0178 \text{ //}$$

\# Acid X reacts with methanol to form ester Y.

$$\begin{matrix} CH_2COOH \\ | \\ CH_2COOH \end{matrix} + 2CH_3OH \rightleftharpoons \begin{matrix} CH_2COOCH_3 \\ | \\ CH_2COOCH_3 \end{matrix} + 2H_2O$$

Acid X　　　　　Ester Y

$\Delta H^\theta = -15$ kJ mol^{-1} //

0

H_2O

$+3Cl_2$

0.25 mol X and 0.34 mol methanol was used to reach eqm in presence of some conc. H_2SO_4. The eqm. mixture has 0.13 mol Y in volume V dm^3.

a) Write an expression for Kc.

$$Kc = \frac{[Y][H_2O]^2}{[X][CH_3OH]^2}$$

b) Calculate no. of moles of X, methanol and water in eqm mixture.

n(X) = 0.25 − 0.13 = 0.12 mol

n(CH_3OH) = $\overset{0.34-}{2}$x 0.13 = ~~0.16 mol~~ 0.18 mol

n(H_2O) = 2x0.13 = 0.16 mol

c) Why is volume not needed to calculate Kc.

→ Volume gets cancelled out in the expression, so its not needed.

d) Calculate Kc with units.

→ $$Kc = \frac{[0.13][0.16]^2}{[0.12][0.18]^2} = 0.856...$$