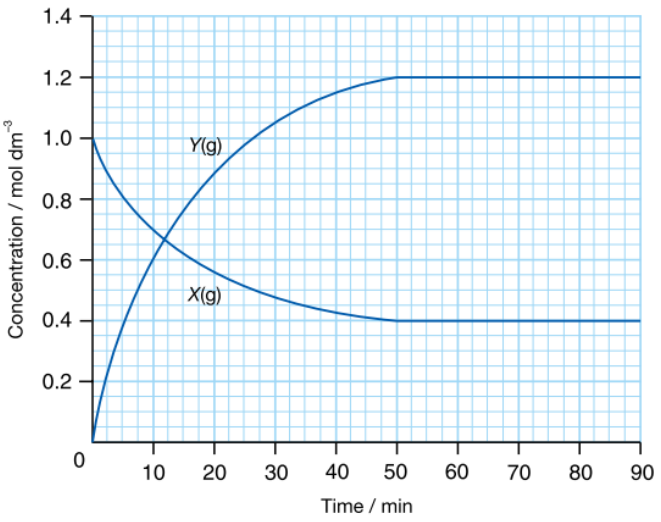
Intensive note (Topic 10: Chemical equilibrium)

Dynamic equilibrium

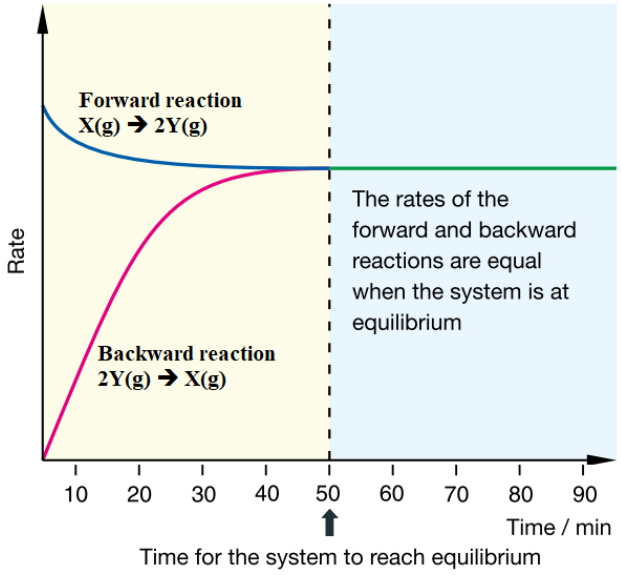
When a chemical system is at dynamic equilibrium, the reactants are converted to the products and the products are converted to the reactants at the same rate. No net change is observed.

Example: Consider the following reaction

$$X(g) \rightleftharpoons 2\,Y(g)$$



Concentration – time graph



Rate – time graph

Characteristics of dynamic equilibrium:

1. **Both reactants and products are present and their concentrations remain unchanged.**
   - In the **concentration-time graph**, both X(g) and Y(g) are present (their concentrations do not drop to 0) and remain unchanged after 50 min.
   - Their equilibrium concentrations do NOT necessarily follow the mole ratio ($[X(g)]_{eqm} : [Y(g)]_{eqm} \neq 1:2$)
   - Their CHANGE in concentration follows the mole ratio (change in [X(g)] : change in [Y(g)] = 0.6:1.2 = 1:2)

2. **The rate of forward reaction is equal to the rate of backward reaction.**
   - In the **rate-time graph**, forward rate and backward rate are the same when the system reached equilibrium (50 min).

3. **Equilibrium can be reached from either direction of a reversible reaction.**

Container	Initial concentration / mol dm^{-3}		Equilibrium concentration / mol dm^{-3}	
	X(g)	Y(g)	X(g)	Y(g)
1	1.0	0.0	0.4	1.2
2	0.0	2.0	0.4	1.2

   - The same equilibrium mixture can be obtained starting with either X(g) (Container 1) or Y(g) (Container 2).

4. **Equilibrium can only be reached in a closed system.**
   - The gaseous mixture has to be kept in a stoppered test tube or a closed container. Otherwise, both X(g) and Y(g) escape and equilibrium can never be reached.

Equilibrium constant calculation

Consider the following reversible reaction:

$$w\,A + x\,B \rightleftharpoons y\,C + z\,D$$

Equilibrium constant (Kc) is a **constant value at a given temperature.** It can be calculated as:

$$K_c = \frac{[C]_{eqm}^{y}[D]_{eqm}^{z}}{[A]_{eqm}^{w}[B]_{eqm}^{x}} \qquad \text{Unit of Kc} = (\text{mol dm}^{-3})^{[(y+z)-(w+x)]}$$

- In the following cases, the term can be omitted and taken to be 1.
  1. The species is a solid
  2. The species is a pure liquid
  3. The species is $H_2O(l)$ and it is a solvent (i.e. other species exist as aqueous state)

The value of Kc indicates the extent of a chemical reaction (how far the reaction proceeds to completion)

- If the value of Kc is large, most of the reactants are converted to products at equilibrium.
  - ➔ The extent of the reaction is very large. The equilibrium position lies in the product side / right.
- If the value of Kc is small, little reactants are converted to products at equilibrium.
  - ➔ The extent of the reaction is very small. The equilibrium position lies in the reactant side / left.
- The magnitude of Kc indicates the extent of a reaction but NOT the rate of the reaction.

Examples

1. Consider the following reversible reaction:

$$A(g) + 2B(g) \rightleftharpoons 2C(g) \qquad \Delta H = -100\ \text{kJ mol}^{-1}$$

A mixture of A(g) and B(g) was introduced into a 2.0 dm^3 closed vessel kept at a fixed temperature. When the system attained equilibrium, the vessel contained 0.2 mol of A(g), 0.1 mol of B(g) and 0.4 mol of C(g)

(a) Calculate the Kc under the conditions of the experiment.

Answer:

$$K_c = \frac{[C(g)]^2}{[A(g)][B(g)]^2}$$

$$K_c = \frac{[\frac{0.4}{2}]^2}{[\frac{0.2}{2}][\frac{0.1}{2}]^2}$$

$$K_c = 160\ \text{mol}^{-1}\text{dm}^3$$

Check the followings:
1. Use concentrations, NOT moles
2. Power
3. Unit of Kc

(b) Consider the following reaction:

$$4C(g) \rightleftharpoons 2A(g) + 4B(g)$$

Calculate the Kc and enthalpy change of the above reaction under the conditions of the experiment.

Answer: $\Delta H = (2)(-1)(-100) = +200\ \text{kJ mol}^{-1}$

$Kc = (160)^{(2)(-1)} = 3.91 \times 10^{-5}\ \text{mol}^2\ \text{dm}^{-6}$

2. Consider the following reversible reaction:

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

In an experiment, 1 mol of $H_2(g)$ and 1 mol of $I_2(g)$ are allowed to react in a closed container at 700 K. When the reaction attains dynamic equilibrium, 1.57 mol of HI(g) is obtained. Calculate Kc for the above reaction at 700 K

Answer:

Mol	$H_2(g)$	$I_2(g)$	HI(g)
Initial	1	1	0
Change	–x	–x	+2x
Equilibrium	1 – x	1 – x	2x

Mole of HI formed = 2x = 1.57

x = 0.785

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

$$K_c = \frac{[\frac{(2)(0.785)}{v}]^2}{[\frac{1-0.785}{v}][\frac{1-0.785}{v}]}$$

$$K_c = 53.3 \text{ (no unit)}$$

- 'Change' row follows the mole ratio (i.e. x : x : 2x).
- Since the no. of mole of HI (product) increased, changes in product side are '+' while changes in reactant side are '–'.
- No. of moles in reactant and no of moles in products are equal. Therefore the volume of container is not necessary while calculating Kc (as all 'v' will be cancelled out during calculation)

3. Consider the reaction represented by the equation below:

$$CH_4(g) + 2H_2S(g) \rightleftharpoons 4H_2(g) + CS_2(g)$$

In an experiment, 1.0 mol of $CH_4(g)$, 2.0 mol of $H_2S(g)$, 4.0 mol of $H_2(g)$ and 1.0 mol of $CS_2(g)$ are mixed in a 50.0 dm^3 closed container maintained at 900 °C. When the reaction attains equilibrium, 6.34 mol of $H_2(g)$ is obtained.

Answer:

Mol	$CH_4(g)$	$H_2S(g)$	$H_2(g)$	$CS_2(g)$
Initial	1.0	2.0	4.0	1.0
Change	–x	–2x	+4x	+x
Equilibrium	1.0 – x	2.0 – 2x	4.0 + 4x	1.0 + x

Mole of H_2 at equilibrium = 4.0 + 4x = 6.34

x = 0.585

$$K_c = \frac{[H_2(g)]^4[CS_2(g)]}{[CH_4(g)][H_2S(g)]^2}$$

$$K_c = \frac{[\frac{4+(4)(0.585)}{50}]^4[\frac{1+(0.585)}{50}]}{[\frac{1-(0.585)}{50}][\frac{2-(2)(0.585)}{50}]^2}$$

$$K_c = 3.58 \text{ mol}^2\text{dm}^{-6}$$

4. Consider the reaction represented by the equation below:

$$2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$$

2.0 mol of NO(g) and 1.0 mol of $O_2(g)$ are allowed to react in a 2.0 dm^3 closed container maintained at 600 K. When the reaction attains equilibrium, 1.62 mol of $NO_2(g)$ is obtained. Calculate the equilibrium constant Kc for the reaction at 600 K.

Answer:

Mol	$NO_2(g)$	NO(g)	$O_2(g)$
Initial	0	2.0	1.0
Change	+2x	–2x	–x
Equilibrium	2x	2 – 2x	1 – x

Mole of NO_2 formed = 2x = 1.62

x = 0.81

$$K_c = \frac{[NO(g)]^2[O_2(g)]}{[NO_2(g)]^2}$$

$$K_c = \frac{[\frac{2-(2)(0.81)}{2}]^2[\frac{1-0.81}{2}]}{[\frac{(2)(0.81)}{2}]^2}$$

$$K_c = 5.23 \times 10^{-3} \text{ mol dm}^3$$

- Equilibrium can be established from either side. Do NOT try to reverse the equation.

5. Consider the following reversible reaction at 298 K:

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$$

In an experiment, 100.0 cm^3 of 0.01 M $Fe_2(SO_4)_3(aq)$ and 100.0 cm^3 of 0.02 M KSCN(aq) were mixed in a conical flask at 298 K, and equilibrium was attained. The concentration of $Fe(SCN)^{2+}(aq)$ in the mixture was 0.007385 M when equilibrium was attained. Calculate the equilibrium constant Kc for the above reaction at 298 K.

Answer: Initial $[Fe^{3+}(aq)] = 0.01 \times (100)/(100 + 100)^* \times 2^{\#} = 0.01$ M

Initial $[SCN^{-}(aq)] = 0.02 \times (100)/(100 + 100)^* = 0.01$ M

Conc.	$Fe^{3+}(aq)$	$SCN^{-}(aq)$	$Fe(SCN)^{2+}(aq)$
Initial	0.01	0.01	0
Change	–x	–x	+x
Equilibrium	0.01 – x	0.01 – x	x

$[Fe(SCN)^{2+}]_{eqm} = x = 0.007385$

$$K_c = \frac{[Fe(SCN)^{2+}(aq)]}{[Fe^{3+}(aq)][SCN^{-}(aq)]}$$

$$Kc = \frac{[0.007385]}{[0.01 - 0.007385][0.01 - 0.007385]}$$

$$Kc = 1080 \text{ mol}^{-1} \text{ dm}^3$$

*Mutual dilution

#1 mole of $Fe_2(SO_4)_3$ gives 2 moles of Fe^{3+}

6. The equilibrium constant Kc for the following reaction is 10.0 at 673 K

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

2.00 mol dm^{-3} of A(g) and 3.00 mol dm^{-3} of B(g) were mixed in a closed container with fixed volume at 673 K. Calculate the equilibrium concentration of A(g) in the system at 673 K.

Answer:

Conc.	A(g)	B(g)	C(g)	D(g)
Initial	2	3	0	0
Change	–x	–x	+x	+x
Equilibrium	2 – x	3 – x	x	x

$$K_c = \frac{[C(g)][D(g)]}{[A(g)][B(g)]}$$

$$10 = \frac{[x][x]}{[2 - x][3 - x]}$$

$$x = 1.75 \text{ or } x = 3.80 \text{ (rejected)}$$

$[A(g)]_{eqm} = 2 - x = 0.247$ mol dm^{-3}

- If there are 2 positive roots, the smaller root should be taken and the larger root should be rejected
- If there is 1 positive root and 1 negative root, the positive root should be taken and the larger root should be rejected

7. Consider the following reversible reaction at 500 K:

$$2X(g) \rightleftharpoons X_2(g)$$

4.0 mol of X(g) is added to a closed container maintained at 500 K. When equilibrium is attained, 20% of X is consumed. Given that the Kc for the above reaction at 500 K is 0.08 mol^{-1} dm^3, calculate the volume of the container.

Answer:

Mol	X(g)	$X_2(g)$
Initial	4.0	0
Change	–2y	+y
Equilibrium	4 – 2y	y

Mole of X consumed = 2y = (4)(20%)

y = 0.4

$$K_c = \frac{[X_2(g)]}{[X(g)]^2}$$

$$0.08 = \frac{[\frac{0.4}{v}]}{[\frac{4 - (2)(0.4)}{v}]^2}$$

$$v = 2.048 \text{ dm}^3$$

Reaction quotient Qc

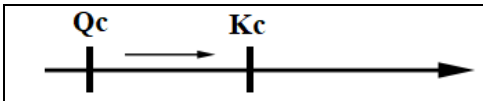
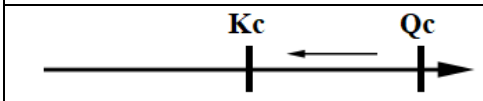
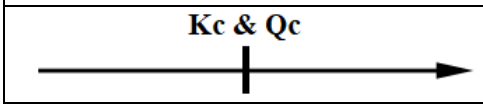
- The direction of shift in the equilibrium position can be predicted by calculating the reaction quotient, Q_c.
- The expression for Q_c is very similar to that of K_c, except that Q_c is calculated from concentrations at any particular moment, not necessarily at equilibrium. Consider the following reversible reaction:

$$w\,A + x\,B \rightleftharpoons y\,C + z\,D$$

The expressions of K_c and Q_c are shown below:

$$K_c = \frac{[C]_{eqm}^{y}[D]_{eqm}^{z}}{[A]_{eqm}^{w}[B]_{eqm}^{x}} \qquad Q_c = \frac{[C]^{y}[D]^{z}}{[A]^{w}[B]^{x}}$$

The equilibrium position can be predicted by comparing Qc with Kc.

	If Qc < Kc, equilibrium position will shift to the right. (Product ↑, reactant ↓) Forward rate will be faster than backward rate.
	If Qc > Kc, equilibrium position will shift to the left. (Reactant ↑, product ↓) Backward rate will be faster than forward rate.
	If Qc = Kc, the system is at equilibrium (reactant and product unchanged) Forward rate and backward rate are equal

Example

At 373 K, the K_c for the following reaction is 0.01 mol dm^{-3}.

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

A 5.00 dm^3 closed container maintained at 373 K initially contains 0.2 mol of A(g), 0.1 mol of B(g) and 0.5 mol of C(g).

(a) Explain whether the concentration of A(g) would increase or decrease just after the reaction started.

Answer:

$$Q_c = \frac{[B(g)][C(g)]}{[A(g)]} = \frac{[\frac{0.1}{5}][\frac{0.5}{5}]}{[\frac{0.2}{5}]}$$

$$Q_c = 0.05 \text{ mol dm}^{-3}$$

As Qc > Kc, equilibrium position will shift to left. [A(g)] would increase just after the reaction started.

(b) Calculate the concentration of A(g) when the system attains equilibrium at 373 K.

Answer:

Mol	A(g)	B(g)	C(g)
Initial	0.2	0.1	0.5
Change	+x	–x	–x
Equilibrium	0.2 + x	0.1 – x	0.5 – x

$$K_c = \frac{[B(g)][C(g)]}{[A(g)]}$$

$$0.01 = \frac{[\frac{0.1-x}{5}][\frac{0.5-x}{5}]}{[\frac{0.2+x}{5}]}$$

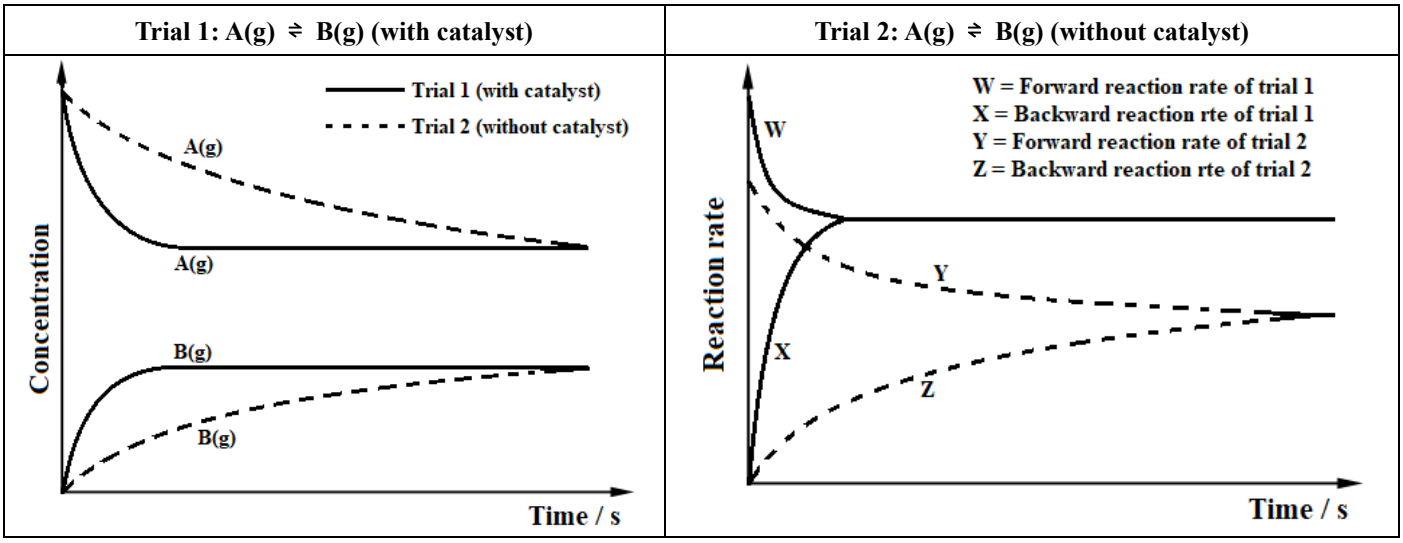
$$x = 0.0688 \; or \; x = 0.581 (rejected)$$

$[A(g)]_{eqm} = 0.2 + 0.0688 = 0.2688$ mol dm^{-3}

Le Châtelier's Principle.

Le Châtelier's Principle states that when a chemical system at equilibrium is disturbed by a change in conditions, the equilibrium position will shift in a direction that tends to counteract the change.

The effect of catalyst on chemical equilibria



Catalyst has no effect has no effect on the equilibrium position. It increases the rates of both forward and backward reactions to the same extent. Catalyst only shortens the time needed to reach equilibrium.

The effect of change in concentration on chemical equilibria

Change	↑ Reactants	↓ Reactants	↑ Products	↓ Products
Response by the system	Consumes reactants Produces products	Produces reactants Consumes products	Consumes products Produces reactants	Produces products Consumes reactants
Eqm position	→	←	←	→
Kc	No change	No change	No change	No change

Example 1

The following equilibrium is established at 298 K.

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$$

Yellow Dark red

(a) Explain what would be observed when a little amount of $Fe(NO_3)_3(s)$ is added to the equilibrium mixture.

Answer: When $Fe(NO_3)_3(s)$ is added, $[Fe^{3+}(aq)]$ increases. (As this stage, Qc < Kc)

The equilibrium position shifts to the right. $[Fe(SCN)^{2+}(aq)]$ increases. Color of the mixture becomes deeper

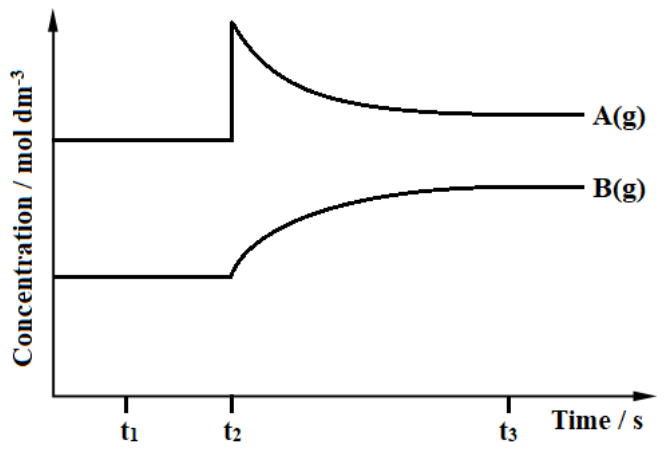
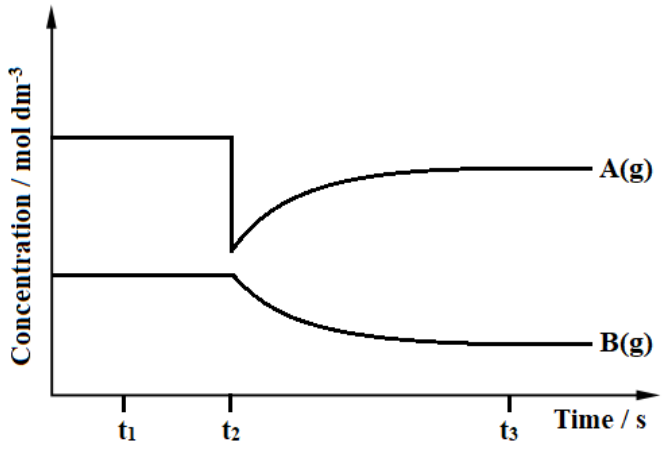
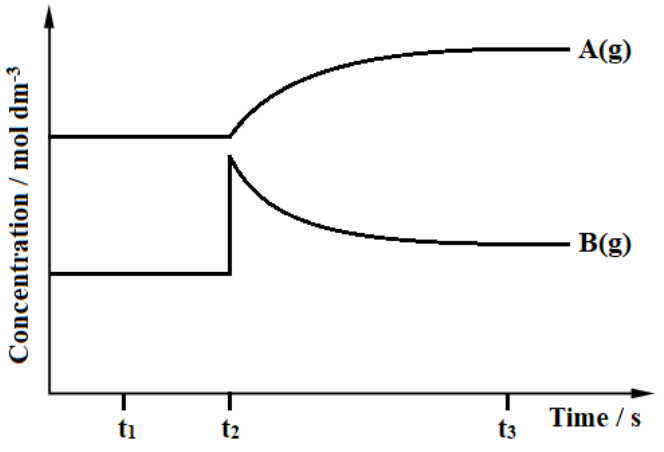
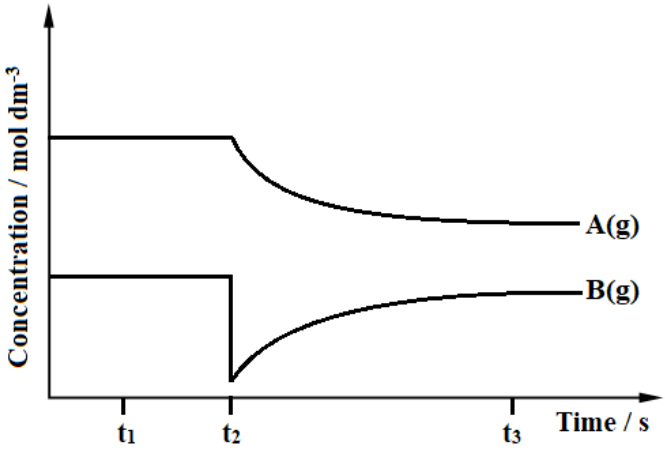
(b) Explain what would be observed when a little amount of $K_2SO_3(s)$ is added to the equilibrium mixture.

Answer: $Fe^{3+}(aq)$ reacts with the $SO_3^{2-}(aq)$. $[Fe^{3+}(aq)]$ decreases. (As this stage, Qc > Kc)

The equilibrium position shifts to the left. $[Fe(SCN)^{2+}(aq)]$ decreases. Color of the mixture becomes paler

Example 2

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

A(g) is added to the equilibrium mixture at t_2	**A(g) is removed to the equilibrium mixture at t_2**
	
- When A(g) is added at t_2, Qc at t_2 < Kc - Equilibrium position will shift to right - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only	- When A(g) is removed at t_2, Qc at t_2 > Kc - Equilibrium position will shift to left - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only
B(g) is added to the equilibrium mixture at t_2	**B(g) is removed to the equilibrium mixture at t_2**
	
- When B(g) is added at t_2, Qc at t_2 > Kc - Equilibrium position will shift to left - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only	- When B(g) is removed at t_2, Qc at t_2 > Kc - Equilibrium position will shift to right - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only

The effect of change in pressure / volume of container on chemical equilibria

Change	No. of moles of gases on reactant > product (➔ = ↓P) (🡐 = ↑P)		No. of moles of gases on reactant = product		No. of moles of gases on product > reactant (➔ = ↑P) (🡐 = ↓P)	
	↑P / ↓V	↓P /↑V	↑P / ↓V	↓P /↑V	↑P / ↓V	↓P /↑V
Response by the system	↓P Produces product	↑P Produces reactant	No effect	No effect	↓P Produces reactant	↑P Produces product
Eqm position	→	←	No change	No change	←	→
Kc	No change	No change	No change	No change	No change	No change
Forward rate	↑	↓	↑	↓	↑	↓
Backward rate	↑	↓	↑	↓	↑	↓

Example 1

The following equilibrium is established at 1000 K in a closed vessel.

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

(a) Explain the effect on the equilibrium position when the pressure of the system is increased at constant temperature.

Answer: The number of moles of gaseous reactant is larger than that of gaseous products.

When the pressure of the system is increased at constant temperature, the equilibrium position shifts to the right

(b) Explain the effect on the equilibrium constant when the pressure of the system is increased at constant temperature.

Answer: Remains unchanged.

As the value of equilibrium constant depends on temperature only.

Example 2

The following equilibrium is established at 700 K in a closed vessel.

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

(a) Explain the effect on the equilibrium position when the volume of the vessel is increased at constant temperature.

Answer: There are equal numbers of moles of gases on the two sides of the equation.

An increase in pressure of the system has no effect on its equilibrium position.

(b) Explain the effect on the forward reaction rate when the volume of the vessel is increased at constant temperature.

Answer: When the volume of the vessel increases, $[H_2(g)]$ and $[I_2(g)]$ decreases.

Therefore the forward reaction rate decreases.

Example 3

The following equilibrium is established at 900 °C in a **x** dm^3 closed vessel.

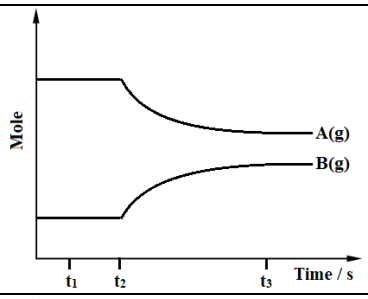
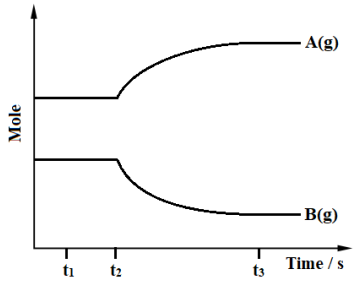
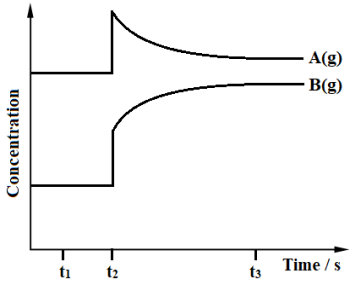
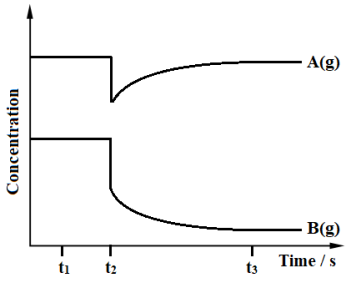
$$CH_4(g) + 2H_2S(g) \rightleftharpoons 4H_2(g) + CS_2(g)$$

When the reaction attains equilibrium, 0.025 mol of $CS_2(g)$ is obtained. The volume of the closed vessel is then changed to **y** dm^3 and a new chemical equilibrium was attained. The number of moles of $CS_2(g)$ was found to be 0.03 in the new chemical equilibrium. Deduce whether **x** or **y** is the larger volume.

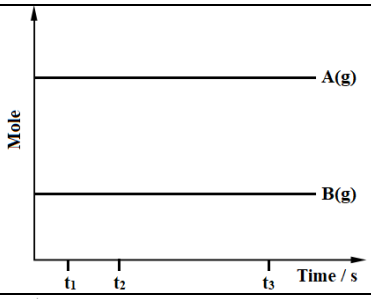
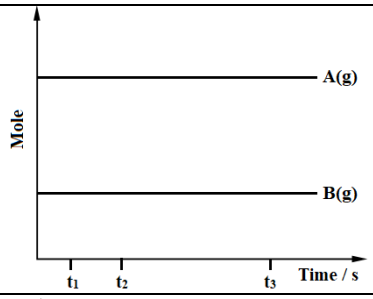
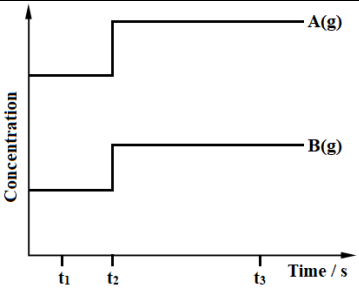
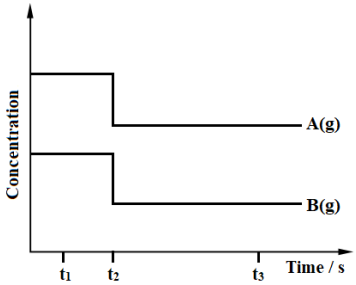
Answer: The no. of moles of $CS_2(g)$ at **y** dm^3 is higher than that at **x** dm^3, that means the equilibrium shifted to the right.

As the number of moles of gaseous reactant is smaller than that of gaseous products, an increase in volume favors the forward reaction. Therefore **y** is the larger volume.

Example 4 $2\,A(g) \rightleftharpoons B(g)$

Volume of container **decreases** at t_2	**Volume** of container **increases** at t_2
	
	
- When volume is decreased at t_2, Qc at t_2 < Kc - Equilibrium position will shift to right - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only	- When volume is increased at t_2, Qc at t_2 > Kc - Equilibrium position will shift to left - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only

Example 5 $A(g) \rightleftharpoons B(g)$

Volume of container **decreases** at t_2	**Volume** of container **increases** at t_2
	
	
- When volume is decreased at t_2, Qc at t_2 = Kc - There is no change in equilibrium position - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only	- When volume is increased at t_2, Qc at t_2 = Kc - There is no change in equilibrium position - The value of Kc at t_1 and t_3 are the same as the value of Kc depends on temperature only

The effect of change in temperature on chemical equilibria

	Forward reaction is endothermic		Forward reaction is exothermic	
	(➔ = endo = ↓T)	(🡐 = exo = ↑T)	(➔ = exo = ↑T)	(🡐 = endo = ↓T)
Change	**↑T**	**↓T**	**↑T**	**↓T**
Response by the system	↓T Produces products	↑T Produces reactants	↓T Produces reactants	↑T Produces products
Eqm position	→	←	←	→
K_c	↑	↓	↓	↑
Forward rate	↑	↓	↑	↓
Backward rate	↑	↓	↑	↓

Example 1

The following equilibrium is established at 298 K.

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$$

Yellow Dark red

(a) It is given that the ΔH of forward reaction is negative. State and explain the expected observable change when the temperature of the mixture is heated to 328 K.

Answer: The forward reaction is exothermic. Increase in temperature will favor endothermic reaction and cause the equilibrium position to shift to the left.

$[Fe(SCN)^{2+}(aq)]$ will decrease and the color of the mixture becomes paler.

(b) When the temperature of the above equilibrium mixture in is lowered, the colour of the solution becomes deeper red. Deduce whether the forward reaction is endothermic or exothermic.

Answer: When the temperature of the mixture decreases, color intensity increases. This indicates that $[Fe(SCN)^{2+}(aq)]$ increases and the equilibrium position shifts to right when the temperature increases.

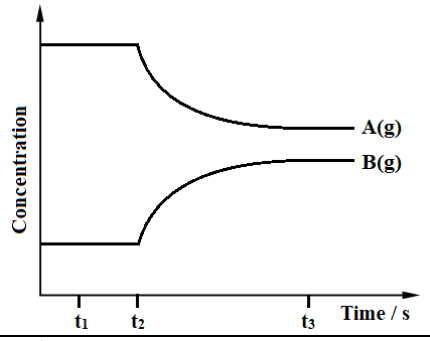
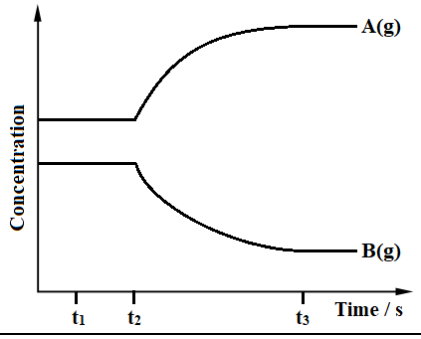
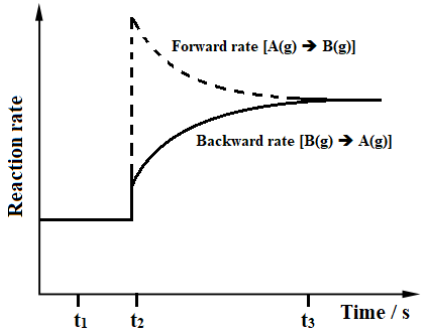
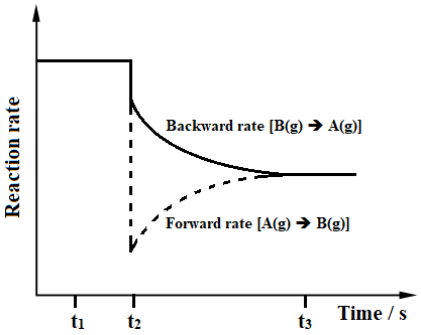
Decrease in temperature favors exothermic reaction. Therefore the forward reaction is exothermic.

(c) I It is given that the ΔH of forward reaction is negative. When the temperature of the equilibrium mixture changed from 298 K to **x** K, a new chemical equilibrium was attained. The color of the mixture becomes paler in the new chemical equilibrium. Deduce whether 298 or **x** is the higher temperature.

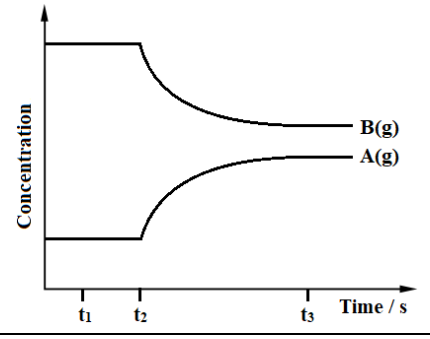
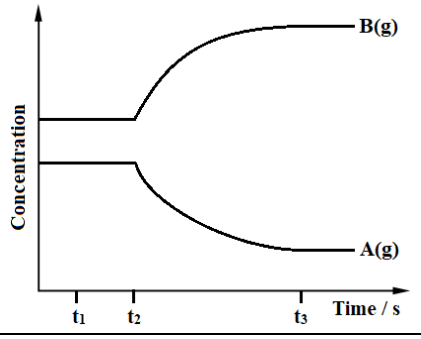
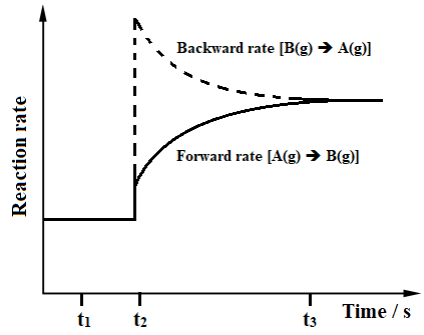
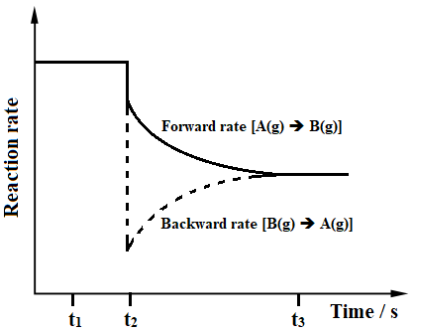
Answer: When the temperature of the mixture changes from 298 K to **x** K, the color intensity decreases. This indicates that $[Fe(SCN)^{2+}(aq)]$ decreases and the equilibrium position shifts to left.

As the backward reaction is endothermic, an increase in temperature favors backward reaction. Therefore **x** K is the higher temperature.

Example 2 $A(g) \rightleftharpoons B(g)$ $\Delta H > 0$

Temperature increases at t_2	**Temperature decreases** at t_2
	
	
- Higher temperature favors endothermic reaction - Equilibrium position will shift to right - K_c at t_3 is greater than K_c at t_1 (Kc increases)	- Lower temperature favors exothermic reaction - Equilibrium position will shift to left - K_c at t_3 is smaller than K_c at t_1 (Kc decreases)

Example 3 $A(g) \rightleftharpoons B(g)$ $\Delta H < 0$

Temperature increases at t_2	**Temperature decreases** at t_2
	
	
- Higher temperature favors endothermic reaction - Equilibrium position will shift to left - K_c at t_3 is smaller than K_c at t_1 (Kc decreases)	- Lower temperature favors exothermic reaction - Equilibrium position will shift to right - K_c at t_3 is greater than K_c at t_1 (Kc increases)

Water-ionization equilibrium

Consider the self-ionization of water at 298 K:

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \qquad \Delta H > 0$$

It is given that at 298 K, $Kc = [H^+(aq)][OH^-(aq)]$. The pH of water equals 7.0 at 298 K.

(a) Calculate the equilibrium constant Kc for the above reaction at 298 K.

Answer: $[H^+(aq)] = 10^{-pH} = 10^{-7}$ mol dm^{-3}

In pure water, $[OH^-(aq)] = [H^+(aq)] = 10^{-7}$ mol dm^{-3}

$Kc = [H^+(aq)][OH^-(aq)] = (10^{-7})(10^{-7}) = 1 \times 10^{-14}$ mol^2 dm^{-6}

(b) Suggest why $[H_2O(l)]$ is considered as a constant with reference to the values of $[H^+(aq)]$ and $[OH^-(aq)]$.

Answer: In 1000 cm^3 water: Mass of H_2O = (Volume of water) x (Density of water) = (1000) x (1) = 1000 g

Mole of H_2O = (Mass) / (Molar mass) = (1000) / (18) = 55.6

Molarity of H_2O = (Mole of water) / (Volume of water) = (55.6) / (1000/1000) = 55.6 M

Since $[H_2O(l)]$ (55.6 M) >>> $[H^+(aq)]$ and $[OH^-(aq)]$ (1×10^{-7} M), $[H_2O(l)]$ does not change and it is regarded as a constant. (The term $[H_2O(l)]$ is omitted in the Kc expression)

(c) Explain whether the pH of water at 328 K would be less than 7.0, equal to 7.0 or greater than 7.0.

Answer: The forward reaction is endothermic. Increase in temperature will favor endothermic reaction and cause the equilibrium position to shift to the right.

$[H^+(aq)]$ will increase and the pH will decrease.

(d) Explain whether water is acidic, neutral or alkaline at 328 K.

Answer: Neutral. At 328 K, $[H^+(aq)]$ and $[OH^-(aq)]$ are equal

*pH of water at 328 K is 6.63. $[H^+(aq)] = 10^{-6.63} = 2.34 \times 10^{-7}$ M. $[OH^-(aq)]$ is also 2.34×10^{-7} M. Since $[H^+(aq)]$ and $[OH^-(aq)]$ are equal, water is still neutral at 328 K.

*Kc at 328 K = $(2.34 \times 10^{-7})(2.34 \times 10^{-7}) = 5.50 \times 10^{-14}$ mol^2 dm^{-6}