

### Le-Chaterlier principle and It's application

61. (c)  $N_2 + O_2 \rightleftharpoons 2NO$ ;  $\Delta n = 0$

62 (a) Increasing the pressure

#### Explanation:

The reaction is **exothermic** (heat is released) and involves a **decrease in volume** (4 moles  $\rightarrow$  2 moles).

By **Le-Chatelier's Principle**, increasing pressure or decreasing temperature shifts equilibrium **toward fewer moles**, i.e., **forward**, forming more ammonia.

63 (b) Low temperature, high pressure

#### Explanation:

The reaction is **exothermic** ( $\Delta H = -ve$ ) and results in **fewer moles of gas** ( $3 \rightarrow 2$ ).

So, **low temperature** (favors exothermic direction) and **high pressure** (favors fewer gas molecules) both push equilibrium **forward**, forming more  $SO_3$ .

64 (c) Increasing the concentration of one or more of the products

#### Explanation:

According to **Le-Chatelier's Principle**, increasing the **product concentration** disturbs the equilibrium.

The system responds by shifting **backward** (toward reactants) to **reduce the excess products** — thus, the **reverse reaction is favoured**.

65. (c) High temperature and low pressure.

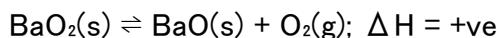
66. (d) High temperature and excess concentration of the reactant concentration.



- 67 (c) Increase in temperature

**Explanation:**

The reaction



is **endothermic**, meaning **heat is absorbed** when it moves forward (producing O<sub>2</sub>).

At equilibrium, **the pressure of O<sub>2</sub>** depends only on **temperature**, not on the amount (mass) of the solids BaO<sub>2</sub> or BaO — because the **concentrations of solids remain constant**.

So, **increasing temperature** shifts the equilibrium **forward**, producing **more O<sub>2</sub>** and thus **increasing its pressure**.

68. (c) Low temperature and high pressure.

69. (a)  $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI} \Rightarrow \Delta n = 2 - 2 = 0$ .

70. (d) Low temperature and low pressure.

- 71 (a) Low pressure

**Explanation:**

For a gaseous reaction,

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

Here,  $\Delta n = (\text{moles of gaseous products}) - (\text{moles of gaseous reactants})$

If  $K_p > K_c$ , then  $(RT)^{\Delta n} > 1$ , so  $\Delta n > 0$ .

That means the number of moles increases in the forward direction.

When the number of moles increases, the reaction is favoured by **low pressure**.

Hence, forward reaction is favoured by **low pressure**.



- 72 (c) It occurs at high temperature

**Explanation:**

Since  $\Delta H$  is positive, the reaction is **endothermic**.

An endothermic reaction proceeds more at **high temperature** according to Le-Chatelier's principle.

Also, the number of gas molecules on both sides is equal ( $2 = 2$ ),  
**so pressure has no effect.**

Therefore, the reaction occurs more at **high temperature**.

73. (c) It is an exothermic reaction hence low temperature and increasing pressure will favour forward reaction

