7.3 — Chemistry JOURNAL —

Fritz Haber: Explore an Equilibrium, Feed the World

SKILLS HANDBOOK A4, A5.1

ABSTRACT

Ammonia gas is produced from elemental nitrogen and oxygen in a reversible reaction, but the equilibrium position heavily favours the reactants at SATP. Fritz Haber conducted basic research on this chemical equilibrium system and fundamentally changed our understanding of equilibria. In a body of investigative work that won him the Nobel Prize in Chemistry in 1916, Haber showed how temperature, pressure, and the addition of a catalyst affected equilibrium position. This work laid the foundation for the chemical engineer Carl Bosch to invent the Haber-Bosch process for industrial-scale ammonia synthesis.

Introduction

The late nineteenth century brought rapid population growth in North America and in Europe, which strained the food supply. To meet the demand, farmers turned to nitrogen-based fertilizers to raise their yields. However, the main sources were naturally occurring and limited: guano (bird droppings), and sodium nitrate from mines. It was known that ammonia and ammonia-containing substances could be used as a fertilizer, but ammonia supplies were also limited. The chemist Fritz Haber (**Figure 1**) was working on improving the energy efficiency of another useful chemical reaction. Impressed by his work, in 1909 the German chemical company BASF hired Haber to investigate the production of ammonia gas from atmospheric nitrogen.

Figure 1 Fritz Haber (1868–1934) made fundamental contributions to the understanding of chemical equilibria.

Investigating Temperature and Pressure

The balanced chemical equation for the production of ammonia gas from its elements is

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

When Haber began his investigations, it was known that this chemical reaction produced very little ammonia gas and that the quantity produced decreased as temperature was increased. Haber's investigations led him to conclude that, when temperature was increased, although the reaction rate increased, the equilibrium position shifted to favour the reactants even more. In his Nobel lecture, Haber explained his next step in his investigations.

> *"To begin with, it was clear that a change to the use of maximum pressure would be advantageous. It would improve the point of equilibrium and probably the rate of reaction as well."*

Unfortunately, Haber found that the temperature of the reaction must be 700 °C to shift the equilibrium position enough to produce greater quantities of ammonia, even at the maximum pressure he could produce with his equipment. The energy needed to create these conditions made this approach uneconomical and so not useful. **Figure 2** shows some of Haber's original data from his work on the effects of temperature and pressure on ammonia production.

t (°C)	T (K)	$\frac{P_{NH_3}}{P_{N_2}{}^{0.5}P_{H2}{}^{0.5}}$	$-\log \frac{P_{NH_3}}{P_{N_2}{}^{0.5}P_{H2}{}^{0.5}}$	Percentage of NH_3 at equilibrium at 1 atm	at 30 atm	at 100 atm	at 200 atm
200	473	0.1807	0.660	15.3	67.6	80.6	85.8
300	573	1.1543	0.070	2.18	31.8	52.1	62.8
400	673	1.8608	0.0138	0.44	10.7	25.1	36.3
500	773	2.3983	0.0040	0.129	3.62	10.4	17.6
600	873	2.8211	0.00151	0.049	1.43	4.47	8.25
700	973	3.1621	0.00069	0.0223	0.66	2.14	4.11
800	1073	3.4417	0.00036	0.0117	0.35	1.15	2.24
900	1173	3.6736	0.000212	0.0069	0.21	0.68	1.34
1000	1273	3.8679	0.000136	0.0044	0.13	0.44	0.87

Figure 2 Data contained in a table prepared by Haber and presented during his Nobel lecture in 1916

An Energy-Efficient Solution

Inspired by the work of other scientists, Haber decided to look for a catalyst that would affect the equilibrium instead. Catalysts do not change the equilibrium position, but they do affect the rate of both the forward and reverse reactions. After much work, Haber discovered that iron oxide could play this role in this equilibrium system, so he was able to produce ammonia at lower temperatures and pressures.

An Application of Insight

Although Haber had achieved his goal and found a way to synthesize ammonia gas from its elements, this did not address the economic and social problems associated with the limited supply of ammonia. Haber discussed this issue during his Nobel lecture:

> "*. . . I was never in doubt that my laboratory work would produce no more than a scientific confirmation of basic principles and a criterion of experimental aids, and that much would need to be added to any success of mine to ensure economic success on an industrial scale. On the other hand, I would hardly have concentrated so much on this problem had I not been convinced of the economic necessity of chemical progress in this field . . . while the immediate object of science lies in its own development, its ultimate aim must be bound up in the moulding influence which it exerts at the right time upon life in general and the whole human arrangement of things around us.*"

Industrial-scale production of ammonia relied on the work of Carl Bosch, BASF's chief chemical engineer. He modified Haber's process and then designed an industrial chemical plant that eventually began producing 10 000 t of ammonia per year. The process is called the Haber–Bosch process to honour the work of both men (**Figure 3**).

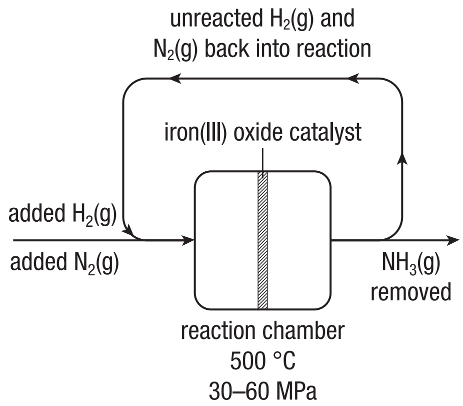

Figure 3 In the Haber–Bosch process, iron(III) oxide is a catalyst and so is not used up. Any reactants present at equilibrium are recycled. The product, ammonia gas, is constantly removed.

Conclusion

Fritz Haber's exploration of the equilibrium system in which ammonia gas is synthesized from its elements was motivated mainly by curiosity. However, the results of his investigations had profound effects on the world outside his laboratory.

Further Reading

Haber, F. (1920). The synthesis of ammonia from its elements. Nobel Lectures, Chemistry 1901–1921. Amsterdam: Elsevier Publishing.

Smil, V. (2004). *Enriching the Earth: Fritz Haber, Carl Bosch, and the Transformation of World Food Production*. Cambridge, MA: MIT Press.

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7.3 Questions

1. Why did the increase in temperature actually lower the ammonia yield? K/U
2. What was the "economic necessity" that Haber discussed in his lecture that encouraged the investigation into the production of ammonia from its elements? K/U
3. Explain how the addition of a catalyst speeds up the production of ammonia. Does the quantity of ammonia at equilibrium change? K/U
4. How does the pressure inside the reaction vessel affect the equilibrium? K/U
5. Both economics and available technologies place limits on the industrial production of chemicals. Describe how economics and the available technologies affected Haber's final process for synthesizing ammonia. How has the process been modified since then? Choose a creative way to communicate your answer. C A
6. Do you think the focus of scientific research should be on solving real-world problems? Why or why not? Refer to the work of Haber in your answer. A

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