In this section of Resonance, we invite readers to pose questions likely to be raised in a classroom situation. We may suggest strategies for dealing with them, or invite responses, or both. “Classroom” is equally a forum for raising broader issues and sharing personal experiences and viewpoints on matters related to teaching and learning science.

Syringe Chemistry

This article presents three experiments that make use of a syringe – a cheap, easily available and normally discarded item – to carry out conventional experiments. They use a minimum of chemicals, little or no heating and reduce the need for costly equipment by reusing old syringes.

The three experiments are:

1. Effect of pressure and temperature on equilibrium between NO₂ and N₂O₄. 2. Effect of pressure on solubility. 3. Ammonia fountain experiment.

Experiment 1. Effect of Pressure and Temperature on Equilibrium

**Aim**: To demonstrate the effect of pressure and temperature changes on the equilibrium involving NO₂ and N₂O₄ by observing change in colour.

**Apparatus**: A 20 cc syringe with needle, conical flask, glass tube, beakers, aluminium foil, Bunsen burner, stands.

**Chemicals**: Pb(NO₃)₂ (lead nitrate), ice.

**Procedure**: The apparatus is set up as shown in Figure 1. A long glass tube is used, so as to keep the ice away
from the fire. Pb(NO$_3$)$_2$ is taken in the conical flask and heated. After a while, when the small beaker is full of brown gas, the syringe is placed at the mouth of the beaker and about 15 cc of the gas is taken in. The needle is removed, and the mouth of the syringe is closed with a finger or a rubber stopper. The effect of pressure is seen by moving the plunger in and out. The pressure is increased and decreased slowly and then rapidly, and the changes in colour observed. Then the syringe is dipped in cold water, and then into hot water, to see the effect of temperature. Again, the changes in the colour of the gas are observed.

**Observation:**

- When the pressure is changed slowly, there is no change in colour.

- When the pressure is increased suddenly, the gas becomes dark brown and then slowly fades back to the original shade.

- When the pressure is decreased suddenly, the gas becomes light brown, and then slowly returns to its original shade.
When the temperature is decreased, the gas becomes light brown, and retains its colour with time. It does not return to its original colour as it did when pressure was changed.

When the temperature is increased, the gas becomes dark brown.

**Explanation:**

Le Chatelier’s Principle: *If a constraint is placed on an equilibrium mixture, then the equilibrium will shift so as to oppose the constraint.*

This is the central idea behind the explanation. The observation for the change in pressure can be explained as follows. According to Le Chatelier’s Principle, the reaction will move in the direction opposing the pressure change. As pressure is directly proportional to number of moles, an increase in pressure will move the reaction towards the side with less moles.

\[
\begin{align*}
\text{brown} & \quad 2\text{NO}_2 & \quad \Rightarrow & \quad \text{colourless} & \quad \text{N}_2\text{O}_4 \\
\text{brown} & \quad \text{the right side has fewer moles}. 
\end{align*}
\]

Thus an increase in pressure favours the forward reaction and vice versa. This can also be seen from the law of equilibria. Let \( P_T \) denote the total pressure. Then, since the equilibrium constant must stay the same:

<table>
<thead>
<tr>
<th>Reaction</th>
<th>( 2\text{NO}_2 )</th>
<th>( \Rightarrow )</th>
<th>( \text{N}_2\text{O}_4 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial moles</td>
<td>2</td>
<td>0</td>
<td></td>
</tr>
<tr>
<td>Moles at equilibrium</td>
<td>( 2(1 - x) )</td>
<td>( x )</td>
<td></td>
</tr>
<tr>
<td>Mole fraction at equilibrium</td>
<td>( 2(1 - x)/(2 - x) )</td>
<td>( x/(2 - x) )</td>
<td></td>
</tr>
<tr>
<td>Partial pressure</td>
<td>( 2(1 - x) \cdot P_T/(2 - x) )</td>
<td>( x \cdot P_T/(2 - x) )</td>
<td></td>
</tr>
</tbody>
</table>
Figure 2. Graph of \( f(x) = \frac{x(2-x)}{4(1-x)^2} \)

Now we compute the equilibrium constant \( k \). As both reactants are gases, concentration equals partial pressure, hence:

\[
k = \frac{[N_2O_4]}{[NO_2]^2} = \frac{(x \cdot P_T) / (2 - x)}{(2(1 - x) \cdot P_T)^2 / (2 - x)^2} = \frac{x(2 - x)}{4(1 - x)^2 \cdot P_T}.
\]

Denoting \( x(2 - x) / 4(1 - x)^2 \) by \( f(x) \), we get:

\[
P_T = \frac{f(x)}{k}, \quad \therefore \quad P_T \propto f(x).
\]

From *Figure 2*, it is clear that \( f(x) \) increases when \( x \) increases. As \( P_T \propto f(x) \), it follows that \( x \) increases as pressure increases; in other words, more \( N_2O_4 \) is formed when pressure increases and the forward reaction is promoted.

But why is there no change in colour when the pressure is changed slowly? This is because if the change is slow, the equilibrium is re-established, and the colour stays the same. However, when the pressure is suddenly increased, the colour gets darker because the same amount of \( NO_2 \) (brown gas) is compressed into a smaller volume. Once the equilibrium is re-established by the formation
of \( N_2O_4 \) (colourless gas) as seen above, the colour returns to its original shade. A similar argument holds for the case when the pressure is decreased suddenly.

For the effect of temperature on equilibrium, we return to Le Chatelier’s Principle:

\[
2\text{NO}_2 \rightleftharpoons \text{N}_2\text{O}_4.
\]

Forward reaction: \( \Delta H \) is negative (\(-13.9\) Kcal) (exothermic).

Thus, if heated, the equilibrium will shift towards \( \text{NO}_2 \), as the backward reaction is endothermic and will absorb heat. Thus, the gas mixture becomes darker. Similarly, if cooled, the forward reaction is promoted and the gas becomes lighter.

**Inference:** Le Chatelier’s principle has been verified in the equilibrium involving \( \text{NO}_2 \) and \( \text{N}_2\text{O}_4 \). As a syringe was used to hold the gas mixture, use of costly equipment to change the pressure of a gas was avoided. Also, minimum quantity of chemicals were used, as only about 15 cc of the gas mixture was needed.

**Experiment 2: Effect of Pressure on Solubility**

**Aim:** To demonstrate the effect of pressure on solubility of \( \text{CO}_2 \) in water.

**Apparatus:** A syringe (20 cc), test tubes, boiling tube.

**Chemicals:** Any carbonate salt, dilute HCl, indicators, methyl orange, universal indicator, phenyl red.

**Procedure:** Ordinary tap water is taken in a test tube and a few drops of the indicator are added to it. \( \text{CO}_2 \) gas is generated by adding dilute HCl to the carbonate salt, and the gas is bubbled through the test tube containing indicator solution. This solution is then taken in the syringe, and the nozzle is kept firmly closed using a finger or a rubber stopper. The plunger is then pulled out to reduce the pressure and the liquid is observed.
**Observation:** As the pressure decreases, bubbles of CO$_2$ form in the liquid and float to the surface.

With methyl orange as the indicator, the colour change is too small to be noticed. However, with universal indicator, the colour changes from reddish orange (pH = 5) to a lighter orange (pH = 5.5) indicating that the solution has become less acidic.

The same procedure was tried for other gases such as NO$_2$ and SO$_2$. However, the evolution of bubbles was not seen with either of these gases.

**Inference:** This indicates that the solubility of CO$_2$ decreases as the pressure is released. The change in pH indicates that CO$_2$ is acidic in aqueous solution.

Once again, the demonstration can be carried out with very small amounts of chemicals using just a syringe and no heating.

**Experiment 3: Ammonia Fountain Experiment**

**Aim:** To replicate the conventional ammonia fountain experiment, using only a syringe.

**Apparatus:** Syringe with a needle, any small dish, plasticine (or any other substance of a similar gummy nature).

**Chemicals:** A bottle of ammonium hydroxide, an indicator (phenol red), water.

**Procedure:** Ammonia gas is sucked into the syringe from the top of the bottle of NH$_4$OH. The nozzle of the syringe is then closed with a wad of plasticine. Water is taken in the dish, and a few drops of the indicator are added. The needle of the syringe is turned around and poked through the wad. The bottom end of the needle is then kept under the indicator solution, as shown in Figure 3. Now the plunger is pulled out slightly to lower the pressure.
**Observation:** The solution is sucked into the syringe with much more force than would be expected.

If the indicator used is phenolphthalein, the solution also turns slightly pink, but no fountain can be observed.

To overcome this, the needle of a 2 cc syringe was used instead of the needle of the 20 cc syringe.

Furthermore, many different indicators were tried. Methyl orange and litmus solution did not change colour. Finally phenol red was chosen. This turned from yellow in plain water to pink in the syringe. Thymol blue also gave a good result, turning from green to blue.

**Inference:** The fact that the solution was sucked in indicates the high solubility of NH₃. The change in colour of the indicator shows the basic nature of NH₃.

The above experiment was repeated with HCl gas taken from the top of a bottle of concentrated HCl. However, the solution was not sucked up very noticeably, and the change in colour did not take place until the syringe was shaken. This might be because HCl is not as highly soluble as NH₃ in water.
Conclusion

The experiments proposed above suggest ways to conduct conventional experiments that minimize the usage of chemicals, heating and laboratory equipment. This approach follows the principles of green chemistry – a concept that aims to make the chemical industry less polluting and energy intensive – on the modest scale of a classroom laboratory. Thus, with some ingenuity and innovation, chemical experiments can be fashioned so as to be better suited for our green future.

Acknowledgement

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Suggested Reading