Adsorption of nickel from aqueous solutions using low cost biowaste adsorbents

Harminder Singh and V. K. Rattan

ABSTRACT

This work has been carried out to check the ability of three biowastes, viz. corn cob ash (CCA), mango stone ash (MSA) and orange peel powder (OPP), to remove Ni(II) ions from aqueous solution. The surface environment of the adsorbents was characterized by Fourier transformation infrared spectroscopy and scanning electron microscopy (SEM) analysis which showed that these adsorbents contain favorable organic groups, such as, amido, amino, hydroxyls, carboxyl groups, etc., on their surface and have uniform characteristics in their surface morphology. The particle sizes for the three screened biowastes were found to be in the range of 1.5–4.0, 0.7–4.5 and 15–35 μm for CCA, MSA and OPP, respectively, as revealed by SEM. The effect of different system variables, viz. adsorbent dose, initial metal ion concentration and pH, were studied. Both Langmuir and Freundlich adsorption models were suitable for describing the sorption of Ni(II) on all three adsorbents used. The maximum sorption capacities of CCA, OPP and MSA used in this study were 107.4, 14.0, and 26.6 mg/g, respectively, for Ni(II) ions at optimum adsorbent dose.

Key words | adsorption, biowastes, Ni(II)

INTRODUCTION

Water is an essential raw material in almost all manufacturing plants, though only a small amount of it may appear in the final product. The remainder becomes a waste material contaminant to a smaller or larger degree, depending on its usage in the plant. After entering the natural water resources it contaminates them (Pandey & Carney 1992).

Toxic metals (Pb, Cu, Cd, Cr, Ni, etc.) are present in the effluents of a variety of industries such as electroplating, leather tanning, cement, mining, dyeing, fertilizer and photography and causes severe environmental and public health problems.

A number of treatment methods for the removal of metal ions from aqueous solutions have been reported, mainly reduction, ion exchange, electrodialysis, electrochemical precipitation, evaporation, solvent extraction, reverse osmosis, chemical precipitation and adsorption (Patterson 1985). Most of these methods suffer from drawbacks such as high capital and operational cost or the disposal of the residual metal sludge.

Adsorption onto solid adsorbents can effectively remove pollutants from both aqueous and gaseous streams and therefore has considerable environmental significance. Activated carbon, the most popular adsorbent, has been traditionally used for the removal of odor, taste and color, which are designated as trace pollutants. Its high adsorptive capacity and versatility have expanded its application to the treatment of numerous industrial waste streams. Other commercial adsorbents, having increased reversibility, have been reviewed (Thomas & Crittenden 1998) and although their versatility and adsorption capacity are generally less than those of activated carbon, they are advantageous for certain applications. Such low cost adsorbents (Adrian et al. 1999; Ho & Mckay 1999a) have found use in laboratory scale for the treatment of various pollutants from water and wastewater.

A number of low-cost adsorbents such as natural neem (Azadirachta indica) sawdust and acid-treated sawdust (Krishnaiah et al. 2006), sphagnum moss peat (Ho &
Mckay 1999b), thai kaolin and ballclay (Chantawong et al. 2003), acid-treated lignin from the pulp and paper industry's black liquor (Pérez et al. 2006), activated biological sludges (Lombrarna et al. 1995; Aksu & Akpınar 2000; Aksu et al. 2002), glacial till soil (Al-Hamdan & Reddy 2006), activated carbon prepared from almond husk (Hasar et al. 2006; Mall et al. 2006; Garg et al. 2008), coir pith (Sudersanan et al. 2006; Thiravetyan et al. 2008), partially converted crab shell waste (Dorris et al. 2005), husk of Lathyrus sativus (Guha et al. 2007), starch, activated charcoal, wood charcoal and clay (Choksi & Joshi 2008), acid-treated rice bran (Zafar et al. 2007), rubber wood ash (Gupta et al. 2000), sawdust (Shukla et al. 2005), tea factory waste (Malkoc & Nuhoglu 2005) and Turkish fly ashes (Bayat 2002) have been used to remove nickel from aqueous solutions and industrial wastewaters.

In the present study, comparison of the adsorption potential of three low cost bio-sorbents, viz. mango stone ash (MSA), orange peel powder (OPP) and corn cob ash (CCA), for the removal of nickel from synthetic aqueous solutions has been checked. The focus of this study is to use the biowastes without any pretreatment so as to make the process economically more viable. Surface properties of these adsorbents have also been checked with the help of Fourier transformation infrared spectroscopy (FTIR) and scanning electron microscopy (SEM).

**MATERIALS AND METHODS**

**Adsorbents preparation**

Adsorbents used in this study are MSA, OPP and CCA, which are waste materials and abundantly available in rural areas of India.

MSA: mango stones were collected from the nearby juice vendors in Chandigarh city. These were washed properly with distilled water and cleaned from any dust and pulp of mango. Then these stones were dried at room temperature. These were then burned in air for 5 h as any other fuel is burned in the rural areas of India, as this study is aimed at the low cost adsorbent preparation and use of any other method can cause an increase in the cost.

OPP: orange peels were collected from the nearby juice vendors as it is considered as a waste by them. They were dried at room temperature and then washed well with distilled water and again dried. Then the orange peels were powdered.

CCA: corn cobs were collected from the nearby villages, dried and washed with distilled water. Then these were burned in air for 5 h.

These adsorbents were stored in an air-tight container for further use. All the above-mentioned adsorbents were sieved through an I.S.70 mesh screen (70 μm) prior to their use.

**Fourier transformation infrared spectroscopy**

It was very important to find out various functional groups present on the surface of the adsorbent, so that possible interactions between the adsorbent and metal ions can be explained. Therefore FTIR spectroscopy was performed for all the adsorbents by using an FTIR spectrophotometer (Perkin Elmer RXI spectrum). KBr discs (150 mg) containing approximately 2% of the adsorbent samples were prepared just before recording the FTIR spectra in the range of 400–4,000 cm⁻¹ and with a resolution of 2 cm⁻¹. The resulting spectra were the average of 16 scans.

**Scanning electron microscopy**

The surface morphology was determined using a JEOL JSM-600 SEM.

**Adsorption studies**

All of the chemical reagents used in this study were of analytical grade. A stock solution of Ni(II) of 1,000 mg/L concentration was prepared by dissolving nickel sulfate in distilled water. All working solutions were prepared by diluting the stock solution with distilled water. The pH adjustments were made using sulfuric acid and sodium hydroxide.

Batch experimental studies were carried out with a known weight of adsorbent and 50 mL working solution of different concentration range (50–500 mg/L) in the 100 mL measuring flasks. Although this concentration
range is much higher than Ni would be in most effluents, high concentrations helped in minimizing the error in the estimation of nickel spectrophotometrically. These flasks were shaken on a wrist shaker at 200 rpm for 1 h. After this, the flasks were kept in desiccators for 24 h with occasional shaking. The basic aim of this study was to check the adsorption potential of biowastes used without any pretreatment and for this purpose only 24 h time was used. This time was sufficient to achieve equilibrium. After equilibration, the adsorbent was separated by filtration through Whatman’s paper no. 40 and the aqueous-phase concentration of the metal ion was analyzed by using standard spectrophotometric methods by producing a more intense color using dimethylglyoxime (detection limit up to 100 μg/50 mL) using a Hitachi 330 UV–VIS spectrophotometer (Basset et al. 1986). The sorption experiments have been carried out in duplicate. Standard deviation and analytical errors are calculated and the maximum error is found to be ±5%. Error bars representing 95% confidence interval are shown in the figures wherever necessary.

The above procedure was repeated with different adsorbent doses (10, 20 and 50 g/L) to check the effect of adsorbent dose. To find the effect of pH on the adsorption of Ni(II), adsorption studies were performed at different pH values in the range of 1–10 taking 1 g each of the adsorbent in 50 mL solutions of 500 mg/L Ni(II) solution at ten different pH values.

The sorption equilibrium \( Q_e \) uptake capacity in mg/g, for each sample is calculated according to mass balance on the metal ion expressed as

\[
Q_e = \left( \frac{C_i - C_e}{m} \right) V
\]  

where \( C_i \) and \( C_e \) are, respectively, initial and equilibrium concentrations of Ni(II), \( m \) is the mass of the adsorbent and \( V \) is the volume of the solution in litres (Krishnaiah et al. 2006).

**Adsorption isotherms**

The adsorption isotherm is a relatively simple method for determining the feasibility of using an adsorbent for a particular application. The liquid phase isotherm shows the distribution of adsorbate between the adsorbed phase and solution phase at equilibrium. It is a plot of the amount of adsorbate adsorbed per unit weight of adsorbent \( Q_e \) versus the concentration of adsorbate remaining in solution \( C_e \).

Adsorption isotherms on different adsorbents are presented as Langmuir and Freundlich isotherms. The former shows initial rapid adsorption tending to be almost constant at higher concentrations.

The linear form of the Langmuir isotherm:

\[
\frac{1}{Q_e} = \frac{1}{Q} + \frac{1}{bQ}C_e
\]

The plots of \( 1/Q_e \) and \( 1/C_e \) for Ni(II) at different adsorbent doses are drawn. From the slope and intercept, the values of constants \( Q \) and \( b \) can be calculated. The constant \( Q \) signifies the adsorption capacity (mg/g) and \( b \) relates to the energy of adsorption (L/mg).

The essential characteristic of the Langmuir equation is expressed in terms of a dimensionless separation factor \( R_L \) (Namasiyayam et al. 2001) and is defined as

\[
R_L = \frac{1}{1 + bC_0}
\]

where \( C_0 \) is the highest initial metal ion concentration (mg/L) and \( b \) is the Langmuir constant. This parameter indicates the isotherm shape according to the following adsorption characteristics: \( R_L > 1 \) unfavorable; \( R_L = 1 \) corresponds to linear; \( 0 < R_L < 1 \) is favorable and \( R_L = 0 \) is irreversible.

The linearized Freundlich isotherm is given by

\[
\log Q_e = \log K_F + \frac{1}{n} \log C_e
\]

Values of \( \log Q_e \) and \( \log C_e \) are plotted for the Ni(II) system to be studied at different adsorbent doses. From the intercept and slope the values of Freundlich constants \( K_F \) and \( n \) can be calculated. The \( K_F \) parameter is relative to adsorption capacity and \( n \) refers to the process intensity (Gupta & Ali 2000). This is frequently used for the interpretation of adsorption from solutions because of its simplicity. Generally, straight-line plots can be obtained by making use of the empirical Freundlich equation.
RESULTS AND DISCUSSION

Characterization of the adsorbents

FTIR of CCA

The FTIR spectrum of CCA (Figure 1(a)) shows a broad absorption band at 3,378.4 cm\(^{-1}\), clearly indicating the presence of \(-\text{OH}\) groups and \(-\text{NH}\) groups. A peak at 2,361.1 is due to C=\(C\) stretch. A sharp peak at 1,653.3 cm\(^{-1}\) confirms the presence of a C=O group. A bend of N–H at 1,559 cm\(^{-1}\) confirms the presence of amide (\(\text{CONH}_2\)) in the adsorbent. Peaks at 1,071.6 and 1,007 cm\(^{-1}\) in the FTIR spectra of CCA can be assigned to the stretching vibrations of C–N and C–O groups, respectively.

FTIR of MSA

In the FTIR spectrum of MSA (Figure 1(b)), a broad absorption band present at 3,377.4 cm\(^{-1}\) reveals the presence of \(-\text{OH}\) and \(-\text{NH}\) groups in MSA. Peaks present at 2,922.7 and 2,365.9 cm\(^{-1}\) are due to the stretching of C–H and C=\(C\) groups, respectively. An absorption peak around 1,568 cm\(^{-1}\) indicates the presence of N–H group. The absorption peaks at 1,115.8 and 1,053.7 cm\(^{-1}\) can be assigned to C–N and C–O stretching, respectively.

FTIR of OPP

The FTIR spectrum (Figure 1(c)) of OPP shows the broad stretching absorption band at 3,390.1 cm\(^{-1}\) revealing the presence of alcoholic and/or phenolic –OH or \(\text{N–H}\) groups (Akar et al. 2005). The band at 2,924.5 cm\(^{-1}\) may be attributed to C–H stretching (Selatnia et al. 2004). The C=O stretch at 1,746.5 cm\(^{-1}\) confirms the presence of an ester group \(\text{C} = \text{O}\) and a strong band at 1,635 cm\(^{-1}\) corresponds to the C=\(C\) stretching. The absorption peaks at 1,519.1 and 1,052 cm\(^{-1}\) can be assigned to N–H bending and C–O stretching, respectively (Pan et al. 2006).

The observations from the FTIR study showed that these adsorbents contain organic groups which favor adsorption on the surface of the adsorbents.

Scanning electron microscopy

SEM micrographs of each material were taken to get a clear view of the grains of the materials. Figures 2(a–c) show scanning electron micrographs of CCA, MSA and OPP, respectively. A close observation of the micrograph of CCA (Figure 2(a)) revealed a structure which is porous in nature. More specifically, it was found that these pores are uniformly distributed throughout the material. The diameter of these pores was found to be in the range of 1.5–4 \(\mu\)m. From Figure 2(b), an idea about the morphology of MSA was assessed and it was found to have some grains with sizes in the range 0.7–4.5 \(\mu\)m. On the other hand, OPP (Figure 2(c)) consists of pellet-like particles distributed uniformly throughout. The particle size for the OPP is in the range of 15–35 \(\mu\)m.

Among these three adsorbents, CCA was found to have a porous character and was expected to perform better than the other two adsorbents.

Effect of adsorbent dose

The dependence of Ni(II) adsorption on these three adsorbents was studied at room temperature and fixed pH values by varying the adsorbent amount from 10 to 30 g/L keeping the adsorbate volume at 50 mL and the Ni(II) concentration at 500 mg/L. The results are shown in Figure 3. The percentage removal of Ni(II) increases rapidly with the increase in the amount of the adsorbents used in this study, which may be due to the greater availability of the exchangeable sites at higher amounts of the adsorbents. However, the amount of Ni(II) adsorbed per gram of the adsorbent was decreased from 18.93 to 13.02 mg/g (for MSA), from 11.69 to 8.19 mg/g (for OPP) and from 44.2 to 16.39 mg/g (for CCA), with increasing adsorbent amount from 10–30 g/L. This decrease in adsorption capacity with increase in the adsorbent dose is mainly because of unsaturation of adsorption sites through the adsorption process (Kovacevic et al. 2000).

Effect of pH

Metal adsorption from aqueous solutions is critically related to pH. It has also been reported that biosorption capacities
for heavy metals are strongly pH sensitive and that adsorption increases as solution pH increases (Zhang et al. 1998; Yu et al. 1999).

To find out the effect of solution pH for the removal efficiency of three different adsorbents, viz. CCA, MSA and OPP, experiments were conducted with metal ion solutions.

Figure 1 | (a)-(c) FTIR spectra of biomass, CCA, MSA and OPP, respectively.
having different pH values. pH of the solution was varied from 1 to 10, keeping all other variables constant, i.e. temperature (25 ± 1°C), 24 h equilibration time, 500 mg/L Ni(II) ion concentration and aqueous to adsorbent ratio 50:1 (mL/g). The effect of pH is shown in Figure 4 in terms of the amount of Ni(II) ions adsorbed per gram (\( Q_e \)) of the adsorbent, respectively, onto the three different adsorbents mentioned above. Initial investigation of adsorption capability (Figure 4) showed that all the adsorbents possessed maximum sorption capacity for the cationic metal ion at or around pH 6. This may be due to the fact that, at these pH values, there was a net negative charge on the surface components of the adsorbents and the ionic state of ligands such as carboxyl, hydroxyl and amino groups will be such as to promote a reaction with metal cations. Overall surface charge on the surface of the adsorbents became positive at lower pH and the presence of H⁺ ions hinders the access...
of metal ions by repulsive forces to the surface functional groups, which resulted in a decrease of the percentage of metal adsorption (Low et al. 1998). Therefore uptake of nickel was less below pH 3, possibly due to the cation competition effects with the oxonium (hydronium) ion H$_3$O$^+$. When the pH of the solution was further increased, soluble hydroxyl species were formed and precipitation of nickel would occur. It is known that at pH 7.7, biosorption of cationic metal decreased, probably because of chemical precipitation. According to the solubility products of metal hydroxide as follows: $K_{sp}(Ni(OH)_2) = 10^{-14}$, solubility studies were meaningless above pH 7.7 due to the formation of insoluble products in the investigated solution (Sing & Yu 1998).

Effect of initial metal ion concentration

The amount of Ni(II) adsorbed per gram of the adsorbent ($Q_e$) for the three chosen adsorbents was observed at different metal ion concentrations (50–500 mg/L). Observation showed that $Q_e$ for all three adsorbents increased with an
An increase in initial metal ion concentration (Figure 5), i.e. from 3.67–18.93, 2.9–11.69 and 4.59–44.2 mg/g for MSA, OPP and CCA, respectively, at adsorbent concentration of 10 g/L. An increase in the initial concentration of metal ions contributes to the driving force to overcome mass transfer resistance of ions between the adsorbent and bulk fluid phases, thus increasing the uptake of metal ions. Moreover, an increase of the concentration of metal ions increases the number of collisions between metal ions and adsorbent, which results in enhancement of the adsorption process (Krishnaiah et al. 2006).

Adsorption isotherms

Heavy metal adsorption data are usually described, analyzed and modeled using an adsorption isotherm, which relates the metal uptake per unit mass of the sorbent to...
the equilibrium metal concentration of the bulk phase (Figures 6(a–c)).

Adsorption study data was also fitted to the Langmuir (Figures 7(a–c)) and Freundlich (Figures 8(a–c)) isotherms and values of various parameters were calculated. Table 1 gives the comparison of monolayer adsorption capacity ($Q$) and energy of adsorption ($b$) and Table 2 gives the comparison of adsorption capacity ($K_F$) and adsorption intensity ($n$). Data fitted well to both isotherms as indicated by the $R$ value which is well above 0.9. But for different materials and at different doses the suitability of these isotherms is different, as shown in Tables 1 and 2 for the Langmuir and Freundlich isotherms, respectively. It was also found that values of $R_L$ were having values between

### Table 2 | Values of different constants for Freundlich isotherms at various doses. The $K_F$ parameter is relative to adsorption capacity and $n$ refers to the process intensity. SD is the standard deviation which represents variation in the values of a variable

<table>
<thead>
<tr>
<th>Parameter/ad sorbent dose in 50 mL</th>
<th>CCA 0.5 (g)</th>
<th>1.0 (g)</th>
<th>1.5 (g)</th>
<th>OPP 0.5 (g)</th>
<th>1.0 (g)</th>
<th>1.5 (g)</th>
<th>MSA 0.5 (g)</th>
<th>1.0 (g)</th>
<th>1.5 (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$R$</td>
<td>0.9966</td>
<td>0.9956</td>
<td>0.9686</td>
<td>0.9734</td>
<td>0.993</td>
<td>0.9926</td>
<td>0.9578</td>
<td>0.997</td>
<td>0.9969</td>
</tr>
<tr>
<td>SD</td>
<td>0.0273</td>
<td>0.0315</td>
<td>0.0618</td>
<td>0.0457</td>
<td>0.0097</td>
<td>0.0314</td>
<td>0.0703</td>
<td>0.0080</td>
<td>0.0253</td>
</tr>
<tr>
<td>$K_F$</td>
<td>1.320</td>
<td>1.288</td>
<td>2.177</td>
<td>0.8943</td>
<td>0.3847</td>
<td>0.4324</td>
<td>1.3564</td>
<td>0.3701</td>
<td>0.03601</td>
</tr>
<tr>
<td>$n$</td>
<td>1.139</td>
<td>1.031</td>
<td>1.119</td>
<td>2.236</td>
<td>1.6790</td>
<td>1.903</td>
<td>2.034</td>
<td>1.283</td>
<td>1.341</td>
</tr>
</tbody>
</table>

### Table 3 | Langmuir adsorption capacity for different low cost adsorbents

<table>
<thead>
<tr>
<th>S. no.</th>
<th>Low cost adsorbent name</th>
<th>Maximum Langmuir adsorption capacity ($Q$ in mg/g)</th>
<th>Reference</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Acid-treated basic lignin from pulp and paper industry's black liquor</td>
<td>5.28</td>
<td>Pérez et al. (2006)</td>
</tr>
<tr>
<td>2</td>
<td>Untreated <em>Chlorella vulgaris</em></td>
<td>264.74</td>
<td>Gaur et al. (2002)</td>
</tr>
<tr>
<td>3</td>
<td>Acid-pretreated <em>C. vulgaris</em></td>
<td>437.90</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>Calcium alginate + <em>C. lutelaTEM05</em></td>
<td>115.99</td>
<td>Ozdemir et al. (2005)</td>
</tr>
<tr>
<td>5</td>
<td>Natural neem sawdust</td>
<td>31.50</td>
<td>Krishnaiah et al. (2006)</td>
</tr>
<tr>
<td>6</td>
<td>Acid-treated neem sawdust</td>
<td>74.10</td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>Free cell of <em>Pseudomonas fluorescens</em> 4F39</td>
<td>145.00</td>
<td>Marqués et al. (2002)</td>
</tr>
<tr>
<td>8</td>
<td>Immobilized cell of <em>Ps. fluorescens</em> 4F39</td>
<td>37.00</td>
<td></td>
</tr>
<tr>
<td>9</td>
<td>Biobeads of cell of <em>Ps. fluorescens</em> 4F39</td>
<td>7.60</td>
<td></td>
</tr>
<tr>
<td>10</td>
<td>Red mud</td>
<td>13.69</td>
<td>Hannachi et al. (2010)</td>
</tr>
<tr>
<td>11</td>
<td>Activated almond husk (MAC-I)</td>
<td>30.77</td>
<td>Hasar (2005)</td>
</tr>
<tr>
<td>12</td>
<td>Sulfuric acid-activated almond husk (MAC-II)</td>
<td>37.17</td>
<td></td>
</tr>
<tr>
<td>13</td>
<td>Dried activated sludge</td>
<td>238.10</td>
<td>Aksu et al. (2002)</td>
</tr>
<tr>
<td>14</td>
<td>Bagasse fly ash</td>
<td>6.19</td>
<td>Mall et al. (2006)</td>
</tr>
<tr>
<td>15</td>
<td>Free biomass of <em>Chlorella sorokiniana</em> (FBCS)</td>
<td>45.87</td>
<td>Akhtar et al. (2004)</td>
</tr>
<tr>
<td>16</td>
<td>Loofa sponge-immobilized biomass of <em>C. sorokiniana</em> (LIBCS)</td>
<td>59.58</td>
<td></td>
</tr>
<tr>
<td>17</td>
<td>Coir pith</td>
<td>15.95</td>
<td>Sudersanan et al. (2006)</td>
</tr>
<tr>
<td>18</td>
<td>Modified coir pith</td>
<td>38.90</td>
<td>Thiravetyan et al. (2008)</td>
</tr>
<tr>
<td>19</td>
<td>Husk of <em>Lathyrus sativus</em> (HLS)</td>
<td>15.70</td>
<td>Guha et al. (2007)</td>
</tr>
<tr>
<td>20</td>
<td>Rice bran</td>
<td>44.90</td>
<td>Zafar et al. (2007)</td>
</tr>
<tr>
<td>21</td>
<td>Rubber wood ash</td>
<td>3.10</td>
<td>Gupta et al. (2000)</td>
</tr>
<tr>
<td>22</td>
<td>Tea factory waste</td>
<td>15.26</td>
<td>Malkoc &amp; Nuhoglu (2005)</td>
</tr>
</tbody>
</table>
0 and 1, which indicates that the adsorption process for these adsorbents was favorable. The Langmuir adsorption capacity (Q) of the three bio-wastes was compared with other low cost adsorbents (Table 3) and found to be comparable with them.

CONCLUSIONS

In this study, sorption of Ni(II) ions on three different bio-wastes, viz. CCA, MSA and OPP, has been investigated. The following conclusions can be drawn:

1. There were certain organic groups present on the surface of the adsorbents, such as, carboxylic, alcoholic, amides, etc, which are known to be favorable for the adsorption process.
2. Surface morphology and surface particle size of all the adsorbents were different from each other, and CCA, having a porous surface, performed better than the other two.
3. pH of the metal ion solution has a significant effect on adsorption and maximum adsorption was found to be around pH 6.
4. The data obtained through this work suggests that all the three adsorbents used in this study are effective low cost biosorbents for the removal of Ni(II) ions in aqueous solution. Although the concentration range (50–500 mg/L) used in this study is much higher than Ni would be in most effluents, high concentrations helped in minimizing the error in the estimation of nickel spectrophotometrically.
5. The sorption of metal ions was also dependent on the amount of adsorbent and concentration of the metal ions.
6. The equilibrium adsorption data were correlated by the Freundlich and Langmuir isotherm equations. The maximum sorption capacities of CCA, OPP and MSA used in this study are 107.4, 14.0 and 26.6 mg/g, respectively, for Ni(II) ions at optimum adsorbent dose.
7. A comparison of the data suggests that CCA is the best biosorbent among the three adsorbents used for this study.
8. The adsorption capacity (Q) for the three adsorbents used in this study is comparable with other low cost adsorbents reported in the literature.

ACKNOWLEDGEMENT

The authors are grateful to the University Grants Commision (UGC), New Delhi, India for providing financial help.

REFERENCES


First received 1 September 2009; accepted in revised form 28 June 2011