Removal of Toxic Mercury (II) from Water via Fe₃O₄/Hydroxyapatite Nanoadsorbent: An Efficient, Economic and Rapid Approach

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Citation

Abstract
A novel nanoadsorbent is synthesized through functionalization of hydroxyapatite with iron oxide (Fe₃O₄/HAP). It is applied as an efficient and economical adsorbent for green removal of toxic mercury (II) ions from water. Crystal structure and morphology of Fe₃O₄/HAP are characterized by XRD and SEM analyses. To its adsorption capabilities, the effect of dosage of adsorbent, contact time, initial pH and initial concentration of mercury ion are scrutinized. To evaluate adsorption behavior, isotherm models and kinetic equation are attended. The experimental data are fitted better to Langmuir isotherm model. The adsorption of mercury on Fe₃O₄/HAP follows pseud-second-order kinetics. Owing to high adsorption capacity of Fe₃O₄/HAP, 492.2 mg/g, Hg²⁺ can be removed considerably, up to 98 %, and the adsorption reaches equilibrium at approximately 10 minutes under optimized conditions. Probing thermodynamic parameters, the adsorption process appears endothermic in nature. Hg (II) adsorption capacity onto Fe₃O₄/HAP increases with enhancing in temperature.

1. Introduction

The common adsorbents primarily include activated carbons, zeolite, clays, biomass and polymeric materials (Crini, 2006; Di Natale et al., 2011; Hsi and Tsai, 2012). However, low adsorption capacities and separation inconvenience are mentioned as these sorbents shortages, which call for new investigation to reach for more efficient materials. Recently few studies have explored the application of mineral materials of environmental functions to dispose wastewater containing heavy metals. Zeolite (Apiratikul and Pavasant, 2008; Pitcher et al., 2004), montmorillonite (Bhattacharyya and Gupta, 2008; Wu et al., 2009), rectorite (Huang et al., 2008a; Huang et al., 2008b), diatomite (Al-Ghouti et al., 2004; Šljivić et al., 2009a) and ion exchange are applied to resolve wastewater issue by adsorption, precipitation and surface complexation. Hydroxyapatite (HAP) (Ca₁₀(PO₄)₆(OH)₂)₃ is a major inorganic component of human bones and teeth (Lim et al., 1997) furthermore it has been considered as a highly efficient co-friendly adsorbent for heavy metal ions, for instance Pb²⁺.
(Kaludjerovic-Radoicic and Raicevic, 2010; Smičiklas et al., 2008), Zn (Kaludjerovic-Radoicic and Raicevic, 2010; Sheha, 2007; Smičiklas et al., 2008), Cu (Kaludjerovic-Radoicic and Raicevic, 2010; Sheha, 2007; Šljivić et al., 2009b; Smičiklas et al., 2008), Cd (Corami et al., 2008; Zhu et al., 2008) and Co (Smičiklas et al., 2006). However, immobilization of heavy metals could lead to HAP application limitation. Therefore, isolation of suspended fine solids from aqueous solution could become a laborious task. In order to improve its applicability for the purification of contaminated water, several composite materials which combine HAP with polymers have been reported (Safavi and Momeni, 2012). Particularly, magnetic Fe₃O₄ has its advantages for the usage as a supportive material since it can be easily manipulated by an external magnetic field. It has a high surface area as well as high efficiency for extraction processes (Chang and Chen, 2005; Dou and Chen, 2011; Najafi et al., 2013). Owing to well-known properties of Fe₃O₄ Nps, several contemporary studies have been carried out on using it for adsorption of heavy metals (Huang and Chen, 2009; Pan et al., 2012). The adsorbents combining magnetic separation technology with adsorption process have been widely used in environmental purification (Pal and Aločijla, 2009; Rocher et al., 2008).

In this study, Fe₃O₄/hydroxyapatite (HAP) composite as a novel nanoadsorbent was prepared to investigate its removal capability of toxic Hg²⁺ from aqueous solutions. Consequently, the impacts of adsorbent dosage, solution pH and competing ions on the adsorption are studied systematically in order to evaluate the efficiency of Fe₃O₄/HAP as a mercury adsorbent. Additionally, the mechanisms of mercury adsorption are discussed according to the isotherms and kinetics models. The Fe₃O₄/HAP demonstrates the remarkable capability comparison among other mercury adsorbents.

2. Materials and Method

2.1. Materials

All solutions are prepared by double distilled de-ionized water. The Hg²⁺ stock solution (100 µg/mL) is prepared readily from HgCl₂ (Merck, Darmstadt, Germany). More diluted mercury solutions are prepared using step-by-step dilution of concentrated mercury solutions. Analytical grades reagent are used including FeCl₃, 4H₂O, FeCl₃, 6H₂O, ammonia (25%), Ca(NO₃)₂, 4H₂O and (NH₄)₂HPO₄.

2.2. Synthesis of Fe₃O₄/HAP

Magnetite nanoparticles (MNps) are synthesized by a co-precipitation method (Safavi and Momeni, 2012). Appropriate amount of FeCl₃. 6H₂O (3.7 mmol) are dissolved in 30 mL of deoxygenated water under a nitrogen atmosphere, at room temperature, and then 10 mL of 25% NH₄OH solution is added to the resulting solution, under vigorous mechanical stirring (200 rpm). A black precipitate is produced instantly. Subsequently, 50 mL of Ca(NO₃)₂. 4H₂O (33.7 mmol) and 50 mL of (NH₄)₂HPO₄ (20 mmol) solutions are prepared. Afterward, ammonia solution with adjusted pH at 11 is dropwise added simultaneously to the solution of obtained precipitate in 30 min with mechanical stirring. The resulting suspension is heated at 90°C for 2 h, then the mixture is cooled down to room temperature and aged for 12-24 h without stirring. The obtained precipitate is separated by a magnet, washed repeatedly with deionized water until neutrality, dried in the drying oven at 90°C, and grinded with mortar. Eventually, Fe₃O₄/HAP is produced as an adsorbent.

2.3. Adsorbent Characterization

The size and morphology of the synthesized Fe₃O₄/HAP are characterized by scanning electron microscopy (SEM) analysis using a LEO 1455 VP microscope (England). Crystal structure of adsorbent is analyzed by X-ray powder diffraction (XRD) pattern recorded on a X-ray diffract meterusing Cu Kα radiation (λ= 0.1541 nm) in steps of 0.05° (2θ) min⁻¹ from 10-80° (2θ).

2.4. Adsorption Experiments

Adsorption isotherm studies are conducted by mixing 0.05 g Fe₃O₄/HAP with 50 mL Hg²⁺ solution in a 250 mL conical flask. For the series of measurements, the ngmL⁻¹ standard, 0.1 mol/L HCl or 0.1 mol/L NaOH solutions are used for pH adjustments. To achieve saturated adsorption, the sample solution is shaken for 10 min at 200 rpm and 25°C. Subsequently, the concentrations of the metal ions are analyzed by cold vapor atomic absorption spectrophotometry (CVAA).

Adsorption kinetic experiments are carried out by batch adsorption method at 25°C on a shaker at 200 rpm. A series of samples are prepared by mixing Hg²⁺ solution (50 mL, 25 ngmL⁻¹) with 0.05 g Fe₃O₄/HAP in a 250 mL conical flask. Standard HCl (0.1 mol/L) or NaOH (0.1 mol/L) solutions are used for pH adjustments. Metal ion concentrations were derived by taking samples at different time intervals. Except for the experiment of pH effect, the pH of the original solution is adjusted at 7. The initial Hg²⁺ concentration of 25 ngmL⁻¹ is used for the adsorption experiments. The adsorption capacity is calculated according to the following equation:

\[ q_e = \frac{V(C_0 - C_e)}{m} \]  

(1)

The removal efficiency of mercury ions is calculated by the difference of mercury ion concentrations in aqueous solution using the following equation:

\[ R(\%) = \frac{C_0 - C_e}{C_0} \times 100 \]  

(2)

Where \( q_e \) (mg/g) is the amount of mercury ions adsorbed onto the unit amount of the adsorbent, \( C_0 \) (mg/L) is the initial mercury ions concentration, \( C_e \) (mg/L) is the final or equilibrium mercury ion concentration, \( V \) (L) is the volume
of the solution, and \( m(g) \) is the adsorbent weight in dry form.

3. Results and Discussion

3.1. Characterization of Prepared Magnetic Adsorbents

The scanning electron microscope (SEM) micrograph shows the synthesized Fe\(_3\)O\(_4\)/HAP with spherical shape and the diameter of about 100 nm (Figure 1). They appear slightly aggregated with Nps, which renders a rough surface and porous structure. Moreover, XRD pattern of the prepared Fe\(_3\)O\(_4\)/HAP adsorbent indicates that the principal components of Fe\(_3\)O\(_4\)/HAP include hydroxyapatite (HAP) and magnetite (Fe\(_3\)O\(_4\)) (Figure 2).

![Figure 1. SEM images Fe\(_3\)O\(_4\)/HAP nanocomposite](image1)

![Figure 2. XRD pattern of the Fe\(_3\)O\(_4\)/HAP nanocomposite.](image2)

3.2. Effect of pH

The influence of pH on the adsorption of mercury ions (50 mL, 25 ngmL\(^{-1}\)) by 50 mg of Fe\(_3\)O\(_4\)/ HAP at 25°C is studied over a pH range of 3-9 (Figure 3). The relationship between the initial pH values and the quantities of Hg\(^{2+}\) ions adsorbed indicate that the Hg\(^{2+}\) ions uptake experiences the continuous increasing by raising the pH value. Moreover, the adsorption amount escalates dramatically in pH range higher than 7. This is attributed to the fact that adsorption Hg\(^{2+}\) ions leads to the reduction of the metal concentration in the aqueous solution at higher pH value.

![Figure 3. The effect of pH of the solution for quantitative removal of Hg\(^{2+}\) ions using Fe\(_3\)O\(_4\)/HAP nanocomposite (metal concentration: 25 ng/mL, absorbent dosage: 50 mg/L, contact time: 10 min, solution volume: 50 ml, rpm: 200); average of three experiments.](image3)

3.3. Effect of Contact Time

The impact of contact time on Hg ions uptake is investigated for duration of 60 min. The results show that adsorption rapidly increases during the first 2 min (Figure 4). Subsequently, after 10min the adsorption reaches an equilibrium state. To obtain adsorption equilibrium for mercury solution, 10 min contact time is needed. Indeed, approximately more than 98% of mercury is adsorbed due to Fe\(_3\)O\(_4\)/HAP high efficiency in this case, and adsorption becomes almost constant up to the end of the experiment. The adsorption is initially rapid, and then slows down considerably. The possible reason might be that Hg (II) ions may easily enter the accessible pore sites and bind with the functional groups in the initial fast adsorption step. The increased number of active sites available at the initial stage might lead to an increased concentration gradient between adsorbate in solution and adsorbate in the adsorbent(Wang et al., 2013). It must be mentioned that for larger solution volumes (> 100 mL) more contact times are needed for complete removal of Hg ions with a fixed amount of Fe\(_3\)O\(_4\)/HAP.

![Figure 4. The effect of contact time on removal of Hg\(^{2+}\) ions using Fe\(_3\)O\(_4\)/HAP (metal concentration: 25 ng/mL, absorbent dosage: 50 mg/L, pH: 7, solution volume: 50 ml, rpm: 200); average of three experiments.](image4)
3.4. Effect of $\text{Fe}_3\text{O}_4$/HAP Dosage

The 60 min uptake experiments are carried out at 25 °C within the $\text{Fe}_3\text{O}_4$/HAP dosage range of 5 to 200 mg. Measurements are based on 50 mL solution (pH 7) with initial Hg(II) ion concentration of 25 ngmL$^{-1}$. It is clearly observed that the removal efficiency is increased as the adsorbent dosage is escalated (Figure 5). When a dosage of 50 mg is used, the removal efficiency reaches approximately 99 %, indicating that the composite adsorbent $\text{Fe}_3\text{O}_4$/HAP shows strong affinity to Hg(II). The adsorption capacity is reached a plateau till the end of the process. In considering the removal efficiency, an adsorption dosage of 50 mg is selected for the next steps.

3.5. Adsorption Isotherm

Adsorption studies over a wide initial concentration range (10-200 ngmL$^{-1}$) are carried to determine the $\text{Fe}_3\text{O}_4$/HAP adsorption capacity for Hg$^{2+}$ ions. The Freundlich and Langmuir adsorption isotherm models are used to fit the experimental data of Hg$^{2+}$ sorption onto $\text{Fe}_3\text{O}_4$/HAP. The linear form of the Freundlich equation can be expressed as follows (Dong et al., 2010; Gong et al., 2009; Ismaiel et al., 2013; Li et al., 2011; Liu et al., 2013):

$$\ln q_e = \ln K_f + \frac{1}{n}\ln C_e$$  \hspace{1cm} (3)

Where $q_e$ is the amount adsorbed at equilibrium concentration (mg/g). $K_f$ is the empirical constant of Freundlich isotherm (L/mg) and $C_e$ is the equilibrium concentration of heavy metal ions in solution (mg/L). The constant $n$ is the empirical parameter related to the intensity of adsorption, which varies with the heterogeneity of the material. When $1/n$ values are in the range 0.1 < $1/n$ < 1, the adsorption process is desirable (Zhang et al., 2007). Langmuir isotherm is originally developed to represent chemisorption on a set of distinct, localized adsorption sites (Asasian et al., 2012). The linear form of the Langmuir isotherm equation is given as the following form (Li et al., 2011; Zhang et al., 2009):

$$\frac{C_e}{q_e} = \frac{1}{K_{eq} q_{max}} + \frac{C_e}{q_{max}}$$ \hspace{1cm} (4)

Where $q_{max}$ is the maximum adsorption capacity (mg/g), and $K_e$ is the Langmuir constant (L/g), which is related to the binding energy of the metal ions to the active site.

Fitting the data using Freundlich adsorption isotherm gives a linear relationship with $R^2 = 0.963$, $K_f = 0.64$ and $1/n = 0.427$, respectively (Figure 6). The linear plot of $C_e/q_e$ against $C_e$ illustrates that the adsorption is in good agreement with the Langmuir model ($R^2 = 0.993$) (Figure 7). The maximum adsorption capacity $q_{max}$ and the Langmuir constant $K_e$ for the adsorption of Hg$^{2+}$ were 0.246 mg/g and 193.79 L/g, respectively. The Langmuir isotherm is valid for monolayer sorption onto a surface with a finite number of identical sites and uniform adsorption energies (Asasian et al., 2012).
3.6. Adsorption Kinetics

The adsorption efficiency is measured as a function of contact time (Figure 8). More than 95% of Hg\textsuperscript{2+} is removed within the first several minutes. A contact time of about 10 min is required to establish the sorption equilibrium of Hg\textsuperscript{2+}. The adsorption kinetics data are analyzed according to the pseudo-second-order kinetics (Freundlich and Hatfield, 1926):

$$\frac{t}{q_t} = \frac{1}{k_2q_e^2} + \frac{t}{q_e}$$  \hspace{1cm} (5)

Where \(k_2\) is the pseudo-second-order rate constant of adsorption (g/mg/min), \(q_e\) and \(q_t\) are the amount of metal ions adsorbed (mg/g) at equilibrium and at time \(t\) (adsorption time, min), respectively. The kinetic process for the adsorption of Hg\textsuperscript{2+} ions by Fe\textsubscript{3}O\textsubscript{4}/HAP excellently follows the characteristics of the pseudo-second-order reaction model \((R^2 = 0.998)\) (Figure 8).

3.7. Adsorption Capacity Comparison

<table>
<thead>
<tr>
<th>Adsorbent</th>
<th>Concentration range</th>
<th>Adsorption time (Min)</th>
<th>Elimination (%)</th>
<th>Ref.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mercaptobenzimidazol-Clay</td>
<td>25-100 mg/L</td>
<td>480</td>
<td>&gt;99%</td>
<td>(Manohar et al., 2002)</td>
</tr>
<tr>
<td>Modified TiO\textsubscript{2}</td>
<td>200 mg/L</td>
<td>32</td>
<td>&gt;99%</td>
<td>(Skubal and Meshkov, 2002)</td>
</tr>
<tr>
<td>coated magnetite</td>
<td>50 µg/L</td>
<td>1200</td>
<td>74%</td>
<td>(Girginova et al., 2010)</td>
</tr>
<tr>
<td>Polyaminil/humic acid</td>
<td>50 mg/L</td>
<td>200</td>
<td>95%</td>
<td>(Zhang et al., 2010)</td>
</tr>
<tr>
<td>TiO\textsubscript{2} Nps</td>
<td>100 mg/L</td>
<td>30</td>
<td>65%</td>
<td>(Dai and Chen, 2011)</td>
</tr>
<tr>
<td>Camel bone charcoal</td>
<td>10 mg/L</td>
<td>30</td>
<td>71%</td>
<td>(Hassan et al., 2008)</td>
</tr>
<tr>
<td>Silver Nps</td>
<td>1 mg/L</td>
<td>1440</td>
<td>68%</td>
<td>(Sumesh et al., 2011)</td>
</tr>
<tr>
<td>Magnetic iron oxide Nps</td>
<td>5-200 µg/L</td>
<td>4</td>
<td>98.6%</td>
<td>(Parham et al., 2012)</td>
</tr>
<tr>
<td>Fe\textsubscript{3}O\textsubscript{4}/HAP</td>
<td>5-200 mg/L</td>
<td>10</td>
<td>99%</td>
<td>This work</td>
</tr>
</tbody>
</table>

The adsorption capacity of adsorbent is determined under optimized conditions (pH7, 25°C, stirring speed= 300 rpm) by batch method. 50 mg adsorbent is added to a 100 mL solution containing 1 mgmL\textsuperscript{-1} of Hg\textsuperscript{2+} ion and stirred for 1 h. Removal percent and adsorbed amount of Hg\textsuperscript{2+} is determined by CV AA of the sample solution before and after removing process. The mercury (II) ion adsorption capacity is determined to be 492.2 mg/g. Comparatively, in most cases, the present nanaosorbent demonstrates a higher adsorption capacity than other adsorbents reported in literatures (Ali, 2012) (Table 1). The higher adsorption capacity for Fe\textsubscript{3}O\textsubscript{4}/HAP is attributed to highest surface area and electrostatic force of attraction between the surface active side and inorganic mercury (II) ions.

3.8. Thermodynamic Study and Temperature Effect

The effects of temperature on mercury (II) adsorption rate on Fe\textsubscript{3}O\textsubscript{4}/HAP is scrutinize at an equilibrium pH of 7 (Figure 9). With the raise in temperature range of 283-313 K the adsorption capacity of Hg (II) increases moderately. This implies that the adsorption is endothermic in nature (Parham et al., 2012). Obviously, adsorption of a solute from solution phase onto solid liquid interface occurs by dislodging water molecules as solvent from interfacial region. To study the thermodynamic properties for adsorption process of mercury (II), the parameters of change in Gibb’s free energy (\(\Delta G^\circ\)), change in enthalpy (\(\Delta H^\circ\)) and change in entropy (\(\Delta S^\circ\)) are obtained by the following equations:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$  \hspace{1cm} (6)

Figure 8. The pseudo-second-order kinetics of Hg\textsuperscript{2+} adsorption on Fe\textsubscript{3}O\textsubscript{4}/HAP (metal concentration: 25 ng/mL, contact time: 10 min, pH: 7, solution volume: 50 ml, rpm: 200).

Figure 9. Effect of temperature on Hg\textsuperscript{2+} adsorption using Fe\textsubscript{3}O\textsubscript{4}/HAP (metal concentration: 25 ng/mL, contact time: 10 min, pH: 7, solution volume: 50 ml, rpm: 200); average of three experiments.
The adsorbent dose (g/L) and the solution interface.

where, m is the adsorbent dose (g/L), q is the amount of mercury (II) adsorbed per unit mass of Fe$_3$O$_4$/HAP(mg/g), $C_e$ is equilibrium concentration (mg/L) and T is temperature in Kelvin. With increasing temperature, the magnitude of $\Delta G^0$ increased, indicating a superior degree of spontaneity at a higher temperature (Table 2). Moreover, due to positive value of $\Delta H^0$, process has endothermic nature. Positive $\Delta S^0$ for mercury (II) exhibited the increased randomness at the adsorbent and solution interface.

### Table 2. Thermodynamic parameters for adsorption of inorganic mercury (II) on Fe$_3$O$_4$/HAP nanocomposite.

<table>
<thead>
<tr>
<th>T (K)</th>
<th>$\Delta G^0$ (kJ/mol)</th>
<th>$\Delta S^0$ (kJ/mol)</th>
<th>$\Delta H^0$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>283</td>
<td>-5.483</td>
<td>0.303</td>
<td>98.80</td>
</tr>
<tr>
<td>293</td>
<td>-7.261</td>
<td>*</td>
<td>*</td>
</tr>
<tr>
<td>303</td>
<td>-9.965</td>
<td>*</td>
<td>*</td>
</tr>
<tr>
<td>313</td>
<td>-10.468</td>
<td>*</td>
<td>*</td>
</tr>
</tbody>
</table>

### 3.9. Effect of Interfering Ions and Solution Volume on Mercury Removal

**Table 3. Effect of some coexistence ions on the removal of Hg$^{2+}$ ions by Fe$_3$O$_4$/HAP (metal concentration: 25 ng/mL, contact time: 10 min, pH: 7, solution volume: 50 ml, rpm: 200); average of three experiments.**

<table>
<thead>
<tr>
<th>Removal cation (%)</th>
<th>Concentration (mg/l)</th>
<th>Compound</th>
<th>Ion type</th>
</tr>
</thead>
<tbody>
<tr>
<td>97.65±2.22</td>
<td>1500</td>
<td>NaCl</td>
<td>Na$^+$</td>
</tr>
<tr>
<td>96.89±4.22</td>
<td>3000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>98.09±5.67</td>
<td>1000</td>
<td>KCl</td>
<td>K$^+$</td>
</tr>
<tr>
<td>98.10±5.54</td>
<td>3000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>95.63±4.43</td>
<td>2000</td>
<td>CaCl$_2$</td>
<td>Ca$^{2+}$</td>
</tr>
<tr>
<td>94.27±2.22</td>
<td>4000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>97.30±4.99</td>
<td>1000</td>
<td>MgSO$_4$</td>
<td>Mg$^{2+}$</td>
</tr>
<tr>
<td>96.71±4.87</td>
<td>4000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>98.10±5.67</td>
<td>1500</td>
<td>NaCl</td>
<td>CI$^-$</td>
</tr>
<tr>
<td>97.20±6.29</td>
<td>3000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>97.72±4.22</td>
<td>1000</td>
<td>MgSO$_4$</td>
<td>SO$_4^{2-}$</td>
</tr>
<tr>
<td>97.30±5.54</td>
<td>2000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>97.17±5.77</td>
<td>1500</td>
<td>KNO$_3$</td>
<td>NO$_3^-$</td>
</tr>
<tr>
<td>96.71±4.91</td>
<td>3000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>98.03±3.78</td>
<td>500</td>
<td>NaHCO$_3$</td>
<td>HCO$_3^-$</td>
</tr>
<tr>
<td>97.61±3.69</td>
<td>1000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>97.16±3.28</td>
<td>500</td>
<td>NaCO$_3$</td>
<td>CO$_3^{2-}$</td>
</tr>
<tr>
<td>97.03±3.60</td>
<td>1000</td>
<td>“</td>
<td>“</td>
</tr>
<tr>
<td>97.22±1.38</td>
<td>500</td>
<td>NH$_4$Cl</td>
<td>NH$_4^+$</td>
</tr>
<tr>
<td>97.34±2.77</td>
<td>1000</td>
<td>“</td>
<td>“</td>
</tr>
</tbody>
</table>

Different feed volumes between 25 and 250 mL are tested. The obtain results (Figure 10) show that the removal of Hg ions are quantitative up to 100 mL of sample volume (removal > 95%). At volumes higher than 100 mL, the analyte is not adsorbed effectively which is probably due to the lower magnetic field strength at higher dilutions (Polshettiwar et al., 2011).

### 4. Conclusion

A fast, simple and low cost magnetic adsorbent for removal of mercury ions from polluted water is successfully developed with a Fe$_3$O$_4$/HAP composite as a novel nanoadsorbent. The results showed that the adsorption efficiency increased with pH increasing and reached a plateau at pH of 7.0. The rapid adsorption and separation of Hg (II) reaches equilibrium within 10 min. At optimized conditions, the adsorption behaviors are fitted better to pseudo-second-order kinetic model and Langmuir isotherm. Fe$_3$O$_4$/HAP revealed high adsorption affinity toward Hg (II) with adsorption capacity of 492.2 mg g$^{-1}$. Thermodynamic study revealed that the adsorption mechanism is endothermic.

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### References


