

contrast, **1** is successfully oxidized in 1–3 M NaOH solution to insoluble and highly stable iridium(IV) hydroxide, $\text{IrO}_2 \cdot n\text{H}_2\text{O}$ (**2**, Scheme 1).^{19,20} The waste (**1**) is processed as a semi-solid or a solution in CH_2Cl_2 to enable a good mass transfer. The reaction proceeds for one day at room temperature and then is driven to completion by heating briefly with added NaClO at 90–100 °C. Reaction completion is detected visually by two key features: (1) black hydroxide **2** is suspended in a yellow solution of $\text{Ir}(\text{OH})_6^{2-}$. In a strongly alkaline media, iridium(IV) exists primarily as a mixture of **2** and $\text{Ir}(\text{OH})_6^{2-}$ with the equilibrium favoring the former. Full dissolution to $\text{Ir}(\text{OH})_6^{2-}$ requires hydrothermal conditions (20 M NaOH, 200 °C),²¹ while full hydrolysis to **2** happens readily at pH 7–9. (2) Adding a fresh portion of NaClO to the completed reaction actively evolves O_2 at 70–100 °C. Once all reducing components are gone, an iridium-catalyzed decomposition of unreacted ClO^- begins.^{22,23} The standard potentials of couples ClO^-/Cl^- (0.890 V) and O_2/OH^- (0.401 V vs. SHE)²⁴ are also consistent with the spontaneous reaction: $2\text{ClO}^- \rightarrow 2\text{Cl}^- + \text{O}_2$.

During the oxidation we observe reversible color changes indicative of Ir(V)/Ir(IV) redox transitions. In a control experiment, a yellow solution of $\text{Ir}(\text{OH})_6^{2-}$ reacts with NaClO at 70 °C turning brown at first, then suddenly evolving O_2 , and turning back to yellow within a few seconds. The sequence can be repeated multiple times suggesting a catalytic behavior of the brown species, apparently iridium(V) (Scheme 1). Although the existence of a transient iridium(V) species has been demonstrated in 2 M HClO_4 ,²⁵ nothing is known about its stability at pH 14. The transient pentavalent iridium could be involved also catalytically in degradation of organic ligands associated with the metal in mixture **1**. The oxidized organics seem to be completely water-soluble, thus facilitating the separation of **2**.

Quantitative precipitation of **2** was achieved by buffering the reaction mixture with excess of solid NaHCO_3 at 100 °C (Scheme 1). During this step, $\text{Ir}(\text{OH})_6^{2-}$ is fully hydrolyzed to **2**. Once the precipitate settles (Fig. S1-C†), a colorless supernatant contains less than 10 ppm iridium, as determined by atomic emission spectroscopy (ICP-OES, Fig. S7†). Sometimes, **2** may form intensely colored colloid solutions because of a slow and incomplete hydrolysis of $\text{Ir}(\text{OH})_6^{2-}$ (Fig. S2†). Fortunately, these colloids can be easily converted to solid **2** by the repeated oxidation (NaClO/NaOH) and buffering steps (NaHCO_3).

Material **2** is amorphous by powder X-ray diffraction. Its iridium content was established gravimetrically by means of $\text{Cs}_2[\text{IrCl}_6]$ as 49.0 wt%, which corresponds to net composition of $\text{IrO}_2 \cdot 9.3\text{H}_2\text{O}$. Unfortunately, the sample cannot be qualified as pure, since it contains up to 2.9 wt% of carbon revealed by the combustion analysis. The carbon content is not changed after treatment with aqueous HCl, suggesting a contamination by non-volatile organics rather than carbonate. Despite the impurity, **2** is easily used in the next steps.

Samples of **2** are resistant to mineral acids, therefore, we utilize a redox approach to convert these to a water-soluble form. A facile dissolution is possible in 6 M HCl in the presence of hydrazine (80 °C, 10 min, Fig. S3-A†). The resulting

solution contains a mixture of iridium(III) aqua-chloride complexes rather than simple IrCl_6^{3-} , since its UV-Vis absorption bands at 415 and 356 nm were not detected.²⁶ Hydrogen peroxide is then added (80 °C, 10 min) to destroy the excess of hydrazine and oxidize Ir(III) to orange-black IrCl_6^{2-} (Fig. S3-B†). Its formation was confirmed by UV-Vis spectroscopy: the characteristic bands were observed at 412, 428, and 486 nm (Fig. S8†).²⁷ Subsequent careful evaporation gives a brown-black crystalline hexachloroiridic acid, $\text{H}_2\text{IrCl}_6 \cdot n\text{H}_2\text{O}$ (**3**, Fig. S3-C†), which was used without purification.²⁸ It is worth noting that excessive heating brings a partial reduction of **3** to Ir(III) chlorides, though not impactful in our case, this process might be detrimental in alternative iridium recycling schemes.

By far, compound **3** is one of the most common precursors for synthetic chemistry involving iridium. Hence, our recycling approach can be utilized to prepare a spectrum of its derivatives. To demonstrate, we synthesized [(1,5-cyclooctadiene) IrCl_2] (**4**) from **3** following a modified reported procedure (Table 1).^{29–31} The reaction involves reduction of **3** in a boiling alcohol solvent in the presence of 1,5-cyclooctadiene ligand. Air-sensitivity of **4** in solutions is well-known,³¹ therefore, the synthesis and isolation were conducted under oxygen-free conditions. Clean **4** was obtained in 87% yield using isopropanol, while the balance of iridium (13%) was recovered as $\text{Cs}_2[\text{IrCl}_6]$ after workup (entry 3). Surprisingly, when $\text{Cs}_2[\text{IrCl}_6]$ is used as a source of iridium in this reaction, only half of it converts to **4** (Table 1, entry 4). The balance persists as beige, insoluble $\text{Cs}_3[\text{IrCl}_6]$ (calc'd 26.5%, found 28.2% Cl). An analytically pure sample of **4** was obtained after slow crystallization from CH_2Cl_2 /hexane (Fig. S4-C†). Finally, the identity of **4** was confirmed by ^1H and $^{13}\text{C}\{^1\text{H}\}$ NMR spectroscopy (Fig. S5 and S6†)³² and the combustion analysis.

In conclusion, we developed a pyrolysis-free process for quantitative iridium recovery from laboratory-generated organoiridium waste that represents the first route available to R&D-scale iridium users. This was essential to our lab, because unlike rhodium, palladium, and ruthenium, we could find no commercial vendor who would recycle our iridium waste. The crucial steps of the process are the oxidative degradation of the waste and the redox dissolution of hydroxide **2**. We show the

Table 1 Synthesis of complex **4**^a

| Entry | Ir source | Alcohol | Yield of 4 , % | Recovered Ir, ^b % |
|-------|--------------------------------------|---------|-----------------------|------------------------------|
| 1 | 2 (0.5 g) → 3 | EtOH | 87 | — |
| 2 | 2 (1.5 g) → 3 | EtOH | 83 | 17 |
| 3 | 2 (1.5 g) → 3 | i-PrOH | 87 | 13 |
| 4 | $\text{Cs}_2[\text{IrCl}_6]$ (3.6 g) | i-PrOH | 49 | 50 |

^a Conditions: **3** or $\text{Cs}_2[\text{IrCl}_6]$ (3.8 mmol), 1,5-cyclooctadiene (6 mL), alcohol (10 mL), and water (10 mL) were refluxed under N_2 for 12 h.
^b Solid leftover materials were treated with hot aqua regia and then with CsCl in a minimum amount of water to form insoluble $\text{Cs}_2[\text{IrCl}_6]$.

practicality of our method by recovering iridium as synthetically important complex **4**. Compounds **2**, **3**, and **4** require minimal to no purification. Because of its low cost and high simplicity, the process enables a very convenient way to reduce metal costs and protect this very precious natural resource. Overall, this route enables a very important option for sustainable use of iridium in laboratory practice and has potential to facilitate iridium recovery on industrial scale.

Author contributions

Conceptualization, data curation, formal analysis, investigation, methodology, project administration, validation, visualization, writing – original draft: V. C. Funding acquisition, resources, supervision, writing – review & editing: T. J. W.

Conflicts of interest

There are no conflicts to declare.

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